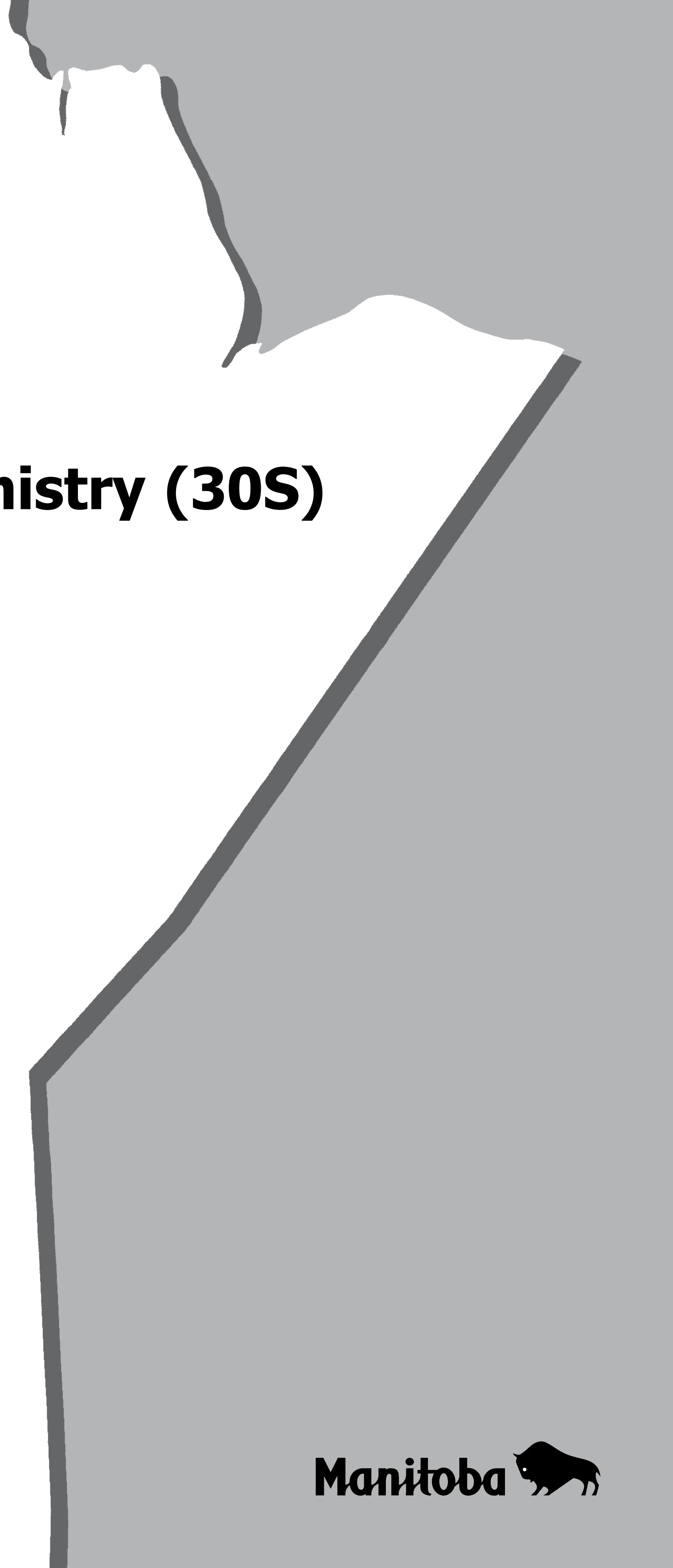
Grade 11 Chemistry (30S)

A Course for
Independent Study



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A Course for Independent Study

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Contents

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Grade 11 Chemistry (30S)

Introduction

Grade 11 Chemistry Introduction

Overview

Welcome to Grade 11 Chemistry. Chemistry is the fascinating study of the interactions between matter and energy. While you have already been studying some chemistry in every science course from Kindergarten to Grade 10, this will be your first full course dedicated to the study of chemistry.

As a student enrolled in an independent study course, you have taken on a dual role—that of a student and a teacher. As a student, you are responsible for mastering the lessons and completing the learning activities and assignments. As a teacher, you are responsible for checking your work carefully, noting areas in which you need to improve and motivating yourself to succeed.

What Will You Learn in This Course?

Not only is chemistry interesting, but it also helps you to understand the world around you and how things work. *Everything* is made up of chemicals, and chemistry helps to explain how everyday things like medication, adhesives (glues), gasoline, concrete, hair colouring, paint, lipstick, and deodorant work. Chemistry explains how soap removes dirt, how food gets cooked faster in a pressure cooker, and why salt is put on icy roads. If you know some chemistry, you can make better choices in your everyday life.

Who Uses Chemistry?

Everybody does. Fire fighters, artists, nurses, truck drivers, doctors, dentists, plumbers, pharmacists, environmentalists, physical therapists, hairdressers, chefs, and veterinarians learn about chemistry and then use that knowledge in their jobs. So, when you start researching different programs of study at universities and colleges, you'll find that you need Grades 11 and 12 Chemistry in order to be admitted to many of them!

How Is This Course Organized?

The Grade 11 Chemistry course consists of the following six modules:

- Module 1: Physical Properties of Matter
- Module 2: Gases and the Atmosphere
- Module 3: Chemical Reactions
- Module 4: Stoichiometry
- Module 5: Solutions
- Module 6: Organic Chemistry

Each module in this course consists of several lessons, which contain the following components:

- **Lesson Focus:** The Lesson Focus at the beginning of each lesson identifies one or more specific learning outcomes (SLOs) that are addressed in the lesson. The SLOs identify the knowledge and skills you should have achieved by the end of the lesson.
- **Introduction:** Each lesson begins by outlining what you will be learning in that lesson.
- **Lesson:** The main body of the lesson consists of the content that you need to learn. It contains information, explanations, diagrams, and completed examples.
- **Learning Activities:** Each lesson has a learning activity that focuses on the lesson content. Your responses to the questions in the learning activities will help you to practise or review what you have just learned. Once you have completed a learning activity, check your responses with those provided in the Learning Activity Answer Key found at the end of the applicable module. Do not send your learning activities to the Distance Learning Unit for assessment.
- **Assignments:** Assignments are found throughout each module within this course. At the end of each module, you will mail or electronically submit all your completed assignments from that module to the Distance Learning Unit for assessment. All assignments combined will be worth a total of 50 percent of your final mark in this course.
- **Summary:** Each lesson ends with a brief review of what you just learned.

This course also includes the following appendices:

- **Appendices A to H:** These are a series of appendices that contain helpful information. As you read through the course, you will be asked to refer to these.

Please note as you read through the course that the definitions of bolded terms may be found in the course glossary (which is Appendix A), and that italicized words represent important information.

What Resources Will You Need for This Course?

You do not need a textbook for this course. All of the content is provided directly within the course.

Required Resources

For this course, you will need access to the following resources. If you do not have access to one or more of these resources, contact your tutor/marker.

- **A molecular model kit:** You will need to have access to this kit in Module 6 in order to construct models of molecules. You can purchase it through the Learning Resource Centre at 1-866-771-6822 or www.manitobalrc.mb.ca. Ask for stock item number 7765.
- **A camera:** You will need access to either a film camera or a digital camera to take pictures of the molecules that you construct in Module 6. You will need to mail or electronically submit these pictures to the Distance Learning Unit.
- **A scientific or graphing calculator:** You will need access to either a scientific calculator or a graphing calculator throughout the course and when writing the midterm and final examinations.
- **A notebook:** You will need a notebook in which to answer the questions from your learning activities.

You will require access to an email account if you plan to

- communicate with your tutor/marker by email
- use the learning management system (LMS) to submit your completed assignments

Optional Resources

It would be helpful if you had access to the following resources:

- **A photocopier/scanner:** With access to a photocopier/scanner, you could make a copy of your assignments before submitting them so that if your tutor/marker wants to discuss an assignment with you over the phone, each of you will have a copy. It would also allow you to continue studying or to complete further lessons while your original work is with the tutor/marker. Photocopying or scanning your assignments will also ensure that you keep a copy in case the originals are lost.
- **A computer with Internet access:** There are many online resources for this course and references are made to them in the lessons where they would be used. However, if you do not have access to a computer, you can still complete the course.

Who Can Help You with This Course?

Taking an independent study course is different from taking a course in a classroom. Instead of relying on the teacher to tell you to complete a learning activity or an assignment, you must tell yourself to be responsible for your learning and for meeting deadlines. There are, however, two people who can help you be successful in this course: your tutor/marker and your learning partner.

Your Tutor/Marker

Tutor/markers are experienced educators who tutor Independent Study Option (ISO) students and mark assignments and examinations. When you are having difficulty with something in this course, contact your tutor/marker, who is there to help you. Your tutor/marker's name and contact information were sent to you with this course. You can also obtain this information in the learning management system (LMS).

Your Learning Partner

A learning partner is someone **you choose** who will help you learn. It may be someone who knows something about chemistry, but it doesn't have to be. A learning partner could be someone else who is taking this course, a teacher, a parent or guardian, a sibling, a friend, or anybody else who can help you. Most importantly, a learning partner should be someone with whom you feel comfortable and who will support you as you work through this course.

Your learning partner can help you keep on schedule with your coursework, read the course with you, check your work, look at and respond to your learning activities, or help you make sense of assignments. You may even study for your examination(s) with your learning partner. If you and your learning partner are taking the same course, however, your assignment work should not be identical.

One of the best ways that your learning partner can help you is by reviewing your midterm and final practice examinations with you. These are found in the learning management system (LMS), along with their answer keys. Your learning partner can administer your practice examination, check your answers with you, and then help you learn the things that you missed.

How Will You Know How Well You Are Learning?

You will know how well you are learning in this course by how well you complete the learning activities, assignments, and examinations.

Learning Activities

The learning activities in this course will help you to review and practise what you have learned in the lessons. You will not submit the completed learning activities to the Distance Learning Unit. Instead, you will complete the learning activities and compare your responses to those provided in the Learning Activity Answer Key found at the end of each module.

Make sure you complete the learning activities. Doing so will not only help you to practise what you have learned, but will also prepare you to complete your assignments and the examinations successfully. Many of the questions on the examinations will be similar to the questions in the learning activities. **Remember that you will not submit learning activities to the Distance Learning Unit.**

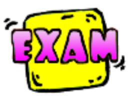
Assignments

Lesson assignments are located throughout the modules and include questions similar to the questions in the learning activities of previous lessons. The assignments have space provided for you to write your answers on the question sheets. You need to show all your steps as you work out your solutions, and make sure your answers are clear (include units, where appropriate).

Once you have completed all the assignments in a module, you will submit them to the Distance Learning Unit for assessment. The assignments are worth a total of 50 percent of your final course mark. You must complete each assignment in order to receive a final mark in this course. You will mail or electronically submit these assignments to the Distance Learning Unit along with the appropriate cover page once you complete each module.

The tutor/marker will mark your assignments and return them to you. Remember to keep all marked assignments until you have finished the course so that you can use them to study for your examinations.

Midterm and Final Examinations



This course contains a midterm examination and a final examination.

- The **midterm examination** is based on Modules 1 to 3, and is worth 20 percent of your final course mark. You will write the midterm examination when you have completed Module 3.
- The **final examination** is based on Modules 1 to 6 and is worth 30 percent of your final course mark. You will write the final examination when you have completed Module 6.

The two examinations are worth a total of **50 percent** of your final course mark. You will write both examinations under supervision.

In order to do well on the examinations, you should review all of the work that you have completed from Modules 1 to 3 for your midterm examination and Modules 1 to 6 for your final examination, including all learning activities and assignments. Please note that 80–85% of the final examination is concentrated on Modules 4 to 6.

You will be required to bring the following supplies when you write both examinations: pens/pencils (2 or 3 of each), an eraser or correction fluid, some blank paper, a ruler, and a scientific or graphing calculator.

Practice Examinations and Answer Keys

To help you succeed in your examinations, you will have an opportunity to complete a Midterm Practice Examination and a Final Practice Examination. These examinations, along with the answer keys, are found in the learning management system (LMS). If you do not have access to the Internet, contact the Distance Learning Unit at 1-800-465-9915 to obtain a copy of the practice examinations.

These practice examinations are similar to the actual examinations you will be writing. The answer keys enable you to check your answers. This will give you the confidence you need to do well on your examinations.

Requesting Your Examinations

You are responsible for making arrangements to have the examinations sent to your proctor from the Distance Learning Unit. Please make arrangements before you finish Module 3 to write the midterm examination. Likewise, you should begin arranging for your final examination before you finish Module 6.

To write your examinations, you need to make the following arrangements:

- **If you are attending school,** your examination will be sent to your school as soon as all the applicable assignments have been submitted. You should make arrangements with your school's ISO school facilitator to determine a date, time, and location to write the examination.
- **If you are not attending school,** check the Examination Request Form for options available to you. Examination Request Forms can be found on the Distance Learning Unit's website, or look for information in the learning management system (LMS). Two weeks before you are ready to write the examination, fill in the Examination Request Form and mail, fax, or email it to

  Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8
Fax: 204-325-1719
Toll-Free Telephone: 1-800-465-9915
Email: distance.learning@gov.mb.ca

How Much Time Will You Need to Complete This Course?

Learning through independent study has several advantages over learning in the classroom. You are in charge of how you learn and you can choose how quickly you will complete the course. You can read as many lessons as you wish in a single session. You do not have to wait for your teacher or classmates.

From the date of your registration, you have a maximum of **12 months** to complete the course, but the pace at which you proceed is up to you. Read the following suggestions on how to pace yourself.

Chart A: Semester 1

If you want to start this course in September and complete it in January, you can follow the timeline suggested below.

Module	Completion Date
Module 1	Middle of September
Module 2	End of September
Module 3	Middle of October
Midterm Examination	Middle of November
Module 4	Beginning of December
Module 5	Middle of December
Module 6	Beginning of January
Final Examination	End of January

Chart B: Semester 2

If you want to start the course in February and complete it in May, you can follow the timeline suggested below.

Module	Completion Date
Module 1	Middle of February
Module 2	Beginning of March
Module 3	Middle of March
Midterm Examination	End of March
Module 4	Middle of April
Module 5	End of April
Module 6	Beginning of May
Final Examination	Middle of May

Chart C: Full School Year (Not Semestered)

If you want to start the course in September and complete it in May, you can follow the timeline suggested below.

Module	Completion Date
Module 1	End of September
Module 2	End of October
Module 3	End of November
Midterm Examination	Beginning of January
Module 4	End of January
Module 5	End of March
Module 6	End of April
Final Examination	Middle of May

Timelines

Do not wait until the last minute to complete your work, since your tutor/marker may not be available to mark it immediately. It may take a few weeks for your tutor/marker to assess your work and return it to you or your school.

> The If you need this course to graduate this school year, all coursework must be received by the Distance Learning Unit on or before the first Friday in May, and all examinations must be received by the Distance Learning Unit on or before the last Friday in May. Any coursework or examinations received after these deadlines may not be processed in time for a June graduation. Assignments or examinations submitted after these recommended deadlines will be processed and marked as they are received.

When and How Will You Submit Completed Assignments?

When to Submit Assignments

While working on this course, you will submit completed assignments to the Distance Learning Unit seven times. The following chart shows you exactly what assignment you will be submitting at the end of each module.

Submission of Assignments	
Submission	Assignments You Will Submit
1	**Module 1: Physical Properties of Matter** Module 1 Cover Sheet 1 Assignment 1.1: Substance Investigation
2	**Module 1: Physical Properties of Matter** Module 1 Cover Sheet 2 Assignment 1.2: Properties of Gases, Liquids, and Solids Assignment 1.3: Dynamic Equilibrium and Phase Changes Assignment 1.4: Vapour Pressure Problems Assignment 1.5: Analyzing a Vapour Pressure Graph
3	**Module 2: Gases and the Atmosphere** Module 2 Cover Sheet Assignment 2.1: Air Quality Improvement Research Assignment 2.2: Determining Significant Digits Assignment 2.3: Solving Problems with Boyle’s Law Assignment 2.4: Investigating the Temperature-Volume Relationship Assignment 2.5: Solving Problems with Charles’ Law Assignment 2.6: Investigating the Temperature-Pressure Relationship Assignment 2.7: Problem Solving with Gay-Lussac’s Law Assignment 2.8: Problem Solving with the Combined Gas Law

continued

Submission of Assignments (continued)	
Submission	Assignments You Will Submit
4	**Module 3: Physical Properties of Matter** Module 3 Cover Sheet Assignment 3.1: Working with Isotopes Assignment 3.2: Working with Chemical Compounds Assignment 3.3: Determining Formula Mass Assignment 3.4: Classifying and Balancing Equations Assignment 3.5: Reaction Types Assignment 3.6: Calculating Molar Mass Assignment 3.7: Determining the Volume of a Gas Assignment 3.8: Converting between Mass, Moles, and Number of Particles Assignment 3.9: Determining Empirical and Molecular Formulas from Percent Composition
5	**Module 4: Stoichiometry** Module 4 Cover Sheet Assignment 4.1: Interpreting a Balanced Equation Assignment 4.2: Using the Molar Ratio Assignment 4.3: Converting between Volume and Mass Assignment 4.4: Solving Limiting Reactant Problems Assignment 4.5: Limiting Reactant Investigation Assignment 4.6: Stoichiometry Applications
6	**Module 5: Solutions** Module 5 Cover Sheet Assignment 5.1: Properties of Solutions Assignment 5.2: The Solution Process Assignment 5.3: Solubility of Polar and Non-Polar Substances Assignment 5.4: Interpreting a Solubility Curve Assignment 5.5: Calculating Solubility Assignment 5.6: Lab Activity: The Effect of Salt on the Melting of Ice Assignment 5.7: Calculations Involving Concentration Assignment 5.8: Determining the Concentration of a Solution Assignment 5.9: Solving Dilution Problems
7	**Module 6: Organic Chemistry** Module 6 Cover Sheet Assignment 6.1: Origins and Major Sources of Hydrocarbons Assignment 6.2: Aliphatic Hydrocarbons Assignment 6.3: Naming and Drawing Alkanes Assignment 6.4: Alkane Isomers* Assignment 6.5: Naming and Drawing Alkenes* Assignment 6.6: Naming and Drawing Alkynes Assignment 6.7: Researching an Aromatic Compound Assignment 6.8: Naming and Drawing Alcohols Assignment 6.9: Drawing Carboxylic Acids Assignment 6.10: Naming and Drawing Esters Assignment 6.11: Chemistry and Our Quality of Life Assignment 6.12: Investigating an Issue in Organic Chemistry * Remember to include the photographs you took in Assignments 6.4, 6.5, and 6.6. You can either mail or electronically submit these photographs.

How to Submit Assignments

In this course, you have the choice of submitting your assignments either by mail or electronically.

- **Mail:** Each time you **mail** something, you must include the print version of the applicable Cover Sheet (found at the end of this Introduction). Complete the information at the top of each Cover Sheet before submitting it along with your assignments.
- **Electronic submission:** You do not need to include a cover sheet when submitting assignments electronically.

Submitting Your Assignments by Mail

If you choose to mail your completed assignments, please photocopy/scan all the materials first so that you will have a copy of your work in case your package goes missing. You will need to place the applicable module Cover Sheet and assignment(s) in an envelope, and address it to

Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Your tutor/marker will mark your work and return it to you by mail.

Submitting Your Assignments Electronically

Assignment submission options vary by course. Sometimes assignments can be submitted electronically and sometimes they must be submitted by mail. Specific instructions on how to submit assignments were sent to you with this course. In addition, this information is available in the learning management system (LMS).

If you are submitting assignments electronically, make sure you have saved copies of them before you send them. That way, you can refer to your assignments when you discuss them with your tutor/marker. Also, if the original hand-in assignments are lost, you are able to resubmit them.

Your tutor/marker will mark your work and return it to you electronically.

The Distance Learning Unit does not provide technical support for hardware-related issues. If troubleshooting is required, consult a professional computer technician.

What Are the Guide Graphics For?

Guide graphics are used through the course to identify and guide you in specific tasks. Each graphic has a specific purpose, as described below.

Assignment: Complete an assignment. You will submit your completed assignments to the Distance Learning Unit for assessment at the specified times.

Learning Activity: Complete a learning activity. This will help you to review or practise what you have learned and to prepare for an assignment or an examination. You will not submit learning activities to the Distance Learning Unit. Instead, you will compare your responses to those provided in the Learning Activity Answer Key found at the end of the applicable module.

Mail or Electronic Submission: Mail or electronically submit your completed assignments to the Distance Learning Unit for assessment at this time.

Learning Partner: Ask your learning partner to help you with this task.

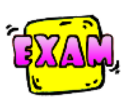
Phone Your Tutor/Marker: Telephone your tutor/marker.

Internet: Use the Internet, if you have access to it, to obtain more information. Internet access is optional for this course.

Check Your Work: Check your responses against those provided in the Learning Activity Answer Key found at the end of the applicable module.

Laboratory Activity: It is time to complete an experiment.

Note: Take note of and remember this important information or reminder.



Examination: Write your midterm or final examination at this time.

Remember: If you have questions or need help at any point during this course, contact your tutor/marker or ask your learning partner for help.

Good luck with the course!

A Note about SLO Numbers

In each lesson introduction, you will find SLO numbers (for example, SLO 1.1.1, SLO 1.1.2, etc.). These numbers have been placed here for teachers. Therefore, if you are a student, you can ignore them.

A Note to Teachers Using This Course as a Classroom Resource

Please note that this course includes three features that have been developed for teachers using this course as a classroom resource:

1. We have written the Specific Learning Outcome (SLO) numbers (SLO 1.1.1, SLO 1.1.2, etc.) in the lesson introductions to indicate which SLO is being taught.
2. We have included a list of the animation and video links referred to in each lesson in Appendix B at the back of the course.
3. We have included a list of the SLOs taught in each lesson in Appendix H at the back of the course.

Grade 11 Chemistry (30S)

Module 1 Cover Sheet 1

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address
Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address
Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Contact Information

Legal Name: ________________________ Preferred Name: ________________________

Phone: ____________________________ Email: ____________________________

Mailing Address: ____

City/Town: ______________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 1 Assignment 1.1	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	____________ Date Received	____________ Date Received
❑ Assignment 1.1: Substance Investigation	________ /17	________ /17
	Total: _____ **/17**	**Total:** _____ **/17**

For Tutor/Marker Use
Remarks:

Grade 11 Chemistry (30S)

Module 1 Cover Sheet 2

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address

Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address

Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Contact Information

Legal Name: ________________________ Preferred Name:________________________

Phone: ____________________________ Email: _______________________________

Mailing Address: ____

City/Town: _______________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 1 Assignments 1.2 to 1.5 Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	**Attempt 1** __________ Date Received	**Attempt 2** __________ Date Received
❑ Assignment 1.2: Properties of Gases, Liquids, and Solids	______ /20	______ /20
❑ Assignment 1.3: Dynamic Equilibrium and Phase Changes	______ /10	______ /10
❑ Assignment 1.4: Vapour Pressure Problems	______ /10	______ /10
❑ Assignment 1.5: Analyzing a Vapour Pressure Graph	______ /20	______ /20
	Total: _____ **/60**	**Total:** _____ **/60**

For Tutor/Marker Use
Remarks:

Grade 11 Chemistry (30S)

Module 2 Cover Sheet

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address
Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address
Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Contact Information

Legal Name: ________________________ Preferred Name: ________________________

Phone: ____________________________ Email: ________________________________

Mailing Address: ____

City/Town: _______________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 2 Assignments	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	_______ Date Received	_______ Date Received
❑ Assignment 2.1: Air Quality Improvement Research	_______ /10	_______ /10
❑ Assignment 2.2: Determining Significant Digits	_______ /10	_______ /10
❑ Assignment 2.3: Solving Problems with Boyle's Law	_______ /21	_______ /21
❑ Assignment 2.4: Investigating the Temperature-Volume Relationship	_______ /20	_______ /20
❑ Assignment 2.5: Solving Problems with Charles' Law	_______ /10	_______ /10
❑ Assignment 2.6: Investigating the Temperature-Pressure Relationship	_______ /10	_______ /10

continued

Grade 11 Chemistry (30S)

Module 2 Cover Sheet *(continued)*

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address

Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address

Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Contact Information

Legal Name: ______________________ Preferred Name: ______________________

Phone: ___________________________ Email: ______________________________

Mailing Address: ____

City/Town: ___________________________________ Postal Code: ____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 2 Assignments *(continued)*	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	______ Date Received	______ Date Received
❑ Assignment 2.7: Problem Solving with Gay Lussac's Law	______ /10	______ /10
❑ Assignment 2.8: Problem Solving with the Combined Gas Law	______ /28	______ /28
	Total: ____ **/119**	**Total:** ____ **/119**

For Tutor/Marker Use
Remarks:

Grade 11 Chemistry (30S)

Module 3 Cover Sheet

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address

Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address

Distance Learning Unit
500–555 Main Street
PO Box 2020
Winkler MB R6W 4B8

Contact Information

Legal Name: ________________________ Preferred Name: ________________________

Phone: ____________________________ Email: ________________________________

Mailing Address: ____

City/Town: _____________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 3 Assignments Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	**Attempt 1** __________ Date Received	**Attempt 2** __________ Date Received
❑ Assignment 3.1: Working with Isotopes	_______ /12	_______ /12
❑ Assignment 3.2: Working with Chemical Compounds	_______ /10	_______ /10
❑ Assignment 3.3: Determining Formula Mass	_______ /10	_______ /10
❑ Assignment 3.4: Classifying and Balancing Equations	_______ /10	_______ /10
❑ Assignment 3.5: Reaction Types	_______ /12	_______ /12
❑ Assignment 3.6: Calculating Molar Mass	_______ /5	_______ /5
❑ Assignment 3.7: Determining the Volume of a Gas	_______ /20	_______ /20

continued

Grade 11 Chemistry (30S)

Module 3 Cover Sheet *(continued)*

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

Drop-off/Courier Address

Distance Learning Unit
555 Main Street
Winkler MB R6W 1C4

Mailing Address

Distance Learning Unit
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Winkler MB R6W 4B8

Contact Information

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Mailing Address: ____

City/Town: _______________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 3 Assignments *(continued)*	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	____________ Date Received	____________ Date Received
❑ Assignment 3.8: Converting between Mass, Moles, and Number of Particles	________ /24	________ /24
❑ Assignment 3.9: Determining Empirical and Molecular Formulas from Percent Compositions	________ /16	________ /16
	Total: _____ **/119**	**Total:** _____ **/119**

For Tutor/Marker Use
Remarks:

Grade 11 Chemistry (30S)

Module 4 Cover Sheet

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

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City/Town: ____________________ Postal Code: ____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	**For Office Use Only**	
Module 4 Assignments	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	________ Date Received	________ Date Received
❑ Assignment 4.1: Interpreting a Balanced Equation	______ /9	______ /9
❑ Assignment 4.2: Using the Molar Ratio	______ /27	______ /27
❑ Assignment 4.3: Converting between Volume and Mass	______ /13	______ /13
❑ Assignment 4.4: Solving Limiting Reactant Problems	______ /24	______ /24
❑ Assignment 4.5: Limiting Reactant Investigation	______ /22	______ /22
❑ Assignment 4.6: Stoichiometry Applications	______ /23	______ /23
	Total: ____ **/118**	**Total:** ____ **/118**

For Tutor/Marker Use
Remarks:

Grade 11 Chemistry (30S)

Module 5 Cover Sheet

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

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Phone: ______________________ Email: ______________________

Mailing Address: ____

City/Town: ______________________ Postal Code: ____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 5 Assignments	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	________ Date Received	________ Date Received
❑ Assignment 5.1: Properties of Solutions	_______ /5	_______ /5
❑ Assignment 5.2: The Solution Process	_______ /10	_______ /10
❑ Assignment 5.3: Solubility of Polar and Non-Polar Substances	_______ /18	_______ /18
❑ Assignment 5.4: Interpreting a Solubility Curve	_______ /10	_______ /10
❑ Assignment 5.5: Calculating Solubility	_______ /14	_______ /14
❑ Assignment 5.6: Lab Activity: The Effect of Salt on the Melting of Ice	_______ /10	_______ /10
❑ Assignment 5.7: Calculations Involving Concentration	_______ /11	_______ /11

continued

Grade 11 Chemistry (30S)

Module 5 Cover Sheet *(continued)*

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

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Mailing Address: ____

City/Town: ____ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 5 Assignments *(continued)*	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	________ Date Received	________ Date Received
❑ Assignment 5.8: Determining the Concentration of a Solution	______ /16	______ /16
❑ Assignment 5.9: Solving Dilution Problems	______ /15	______ /15
	Total: ____ **/109**	**Total:** ____ **/109**

For Tutor/Marker Use
Remarks:

GRADE 11 CHEMISTRY (30S)

Module 6 Cover Sheet

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

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Contact Information

Legal Name: ________________________ Preferred Name: ________________________

Phone: ____________________________ Email: ______________________________

Mailing Address: ____

City/Town: ______________________________________ Postal Code: _____________

Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 6 Assignments Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	**Attempt 1** _______ Date Received	**Attempt 2** _______ Date Received
❑ Assignment 6.1: Origins and Major Sources of Hydrocarbons	_______ /6	_______ /6
❑ Assignment 6.2: Aliphatic Hydrocarbons	_______ /5	_______ /5
❑ Assignment 6.3: Naming and Drawing Alkanes	_______ /15	_______ /15
❑ Assignment 6.4: Alkane Isomers (include photographs)	_______ /15	_______ /15
❑ Assignment 6.5: Naming and Drawing Alkenes (include photographs)	_______ /16	_______ /16
❑ Assignment 6.6: Naming and Drawing Alkynes (include photographs)	_______ /15	_______ /15
❑ Assignment 6.7: Researching an Aromatic Compound	_______ /12	_______ /12
❑ Assignment 6.8: Naming and Drawing Alcohols	_______ /10	_______ /10
❑ Assignment 6.9: Drawing Carboxylic Acids	_______ /5	_______ /5

continued

Grade 11 Chemistry (30S)

Module 6 Cover Sheet *(continued)*

Please complete this sheet and place it on top of your assignments to assist in proper recording of your work. Submit the package to:

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Attending School: ❑ No ❑ Yes

School Name: ____

Has your contact information changed since you registered for this course? ❑ No ❑ Yes

Note: Please keep a copy of your assignments so that you can refer to them when you discuss them with your tutor/marker.

For Student Use	For Office Use Only	
Module 6 Assignments *(continued)*	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	_______________ Date Received	_______________ Date Received
❑ Assignment 6.10: Naming and Drawing Esters	________ /16	________ /16
❑ Assignment 6.11: Chemistry and Our Quality of Life	________ /10	________ /10
❑ Assignment 6.12: Investigating an Issue in Organic Chemistry	________ /16	________ /16
	Total: _____ **/141**	**Total:** _____ **/141**

For Tutor/Marker Use
Remarks:

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Grade 11 Chemistry (30S)

Module 1: Physical Properties of Matter

Module 1: Physical Properties of Matter

Introduction

Have you ever wondered what is really going on inside a plasma television? How does the evaporation of a liquid in a closed container compare with that of a liquid in an open container? Maybe you have seen the label on a tank of compressed gas (like the one attached to your barbeque) and wondered why it matters how the tank of gas is stored. As you work through these five lessons, these (and many other) questions will be answered.

In this module, you will learn about the four states of matter and how to describe the size, motion, and energy of particles in each. You will revisit phase changes and find out how particles behave during these changes, as well as the role kinetic energy plays in the process. Next, you will focus specifically on one phase change, evaporation, and variables that affect the vapour pressure created by this process. Finally, you will put your graphing skills to use when you learn how to plot and interpret a Vapour Pressure Curve.

General Notes

Here are a few items to be aware of as you work through this course:

- Assignments and learning activities are numbered sequentially. This means that their numbers will not always be the same as the lesson number.
- Not every lesson has both an assignment and a learning activity.
- The timelines given in the Introduction of each lesson are only guidelines; the lesson may take you less time or more time.

From time to time you will see a text box in the margin. These text boxes contain interesting information that pertains to the lesson. You will not be assessed on this information, which means you will not be asked about it in learning activities, assignments, or examinations.

Assignments in Module 1

When you have completed the assignments for Module 1, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 1.1	Substance Investigation
2	Assignment 1.2	Properties of Gases, Liquids, and Solids
3	Assignment 1.3	Dynamic Equilibrium and Phase Changes
4	Assignment 1.4	Vapour Pressure Problems
5	Assignment 1.5	Analyzing a Vapour Pressure Graph

Once you have completed Assignment 1.1, you will immediately submit that work to the Distance Learning Unit. You will wait until you have worked through the rest of Module 1 before you submit Assignments 1.2, 1.3, 1.4, and 1.5. The instructions for submitting assignments are found in the course Introduction.

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

Lesson 1: States of Matter (2 hours)

Lesson Focus

SLO C11-1-01: Describe the properties of gases, liquids, solids, and plasma.
Include: density, compressibility, diffusion

Lesson Introduction

Since Grade 7 you have been continuing your exploration of several chemical principles that will help you in this course. You have already learned something about chemical and physical properties, as well as the states of matter. For example, water, ice, and steam are all words that can be used to describe water, depending on the circumstances. In this lesson, you will take things a step further and describe some specific chemical properties of solids, liquids, gases, and plasma. Learning more about the properties of matter can help us to identify unknown substances and use them appropriately.

Chemical and Physical Properties

Let's assume that everyone needs a bit of review before we go any further. There are some important terms that you need to know in order to understand this lesson and the ones that follow.

When we discuss terms like gases, liquids, solids, and plasma, we are really talking about matter. Do you remember what matter is? Why does it matter? **Matter** is anything that takes up space (has a volume) and has a mass. In other words, everything that makes up our physical and chemical world is matter.

A **property** is a characteristic that we can use to help us identify a person, place, or thing. This is not only a term used in chemistry, but one that is used commonly in other areas. Describing how your room looks requires referring to the properties of your room, such as blue carpet and red walls. In chemistry, there are two types of properties we use—chemical properties and physical properties.

Chemical Properties

Chemical properties describe how a substance reacts with other substances. As the substance changes, these types of properties can be observed. Examples include rusting and creating gas bubbles. Remember that chemical properties can be either **qualitative** (descriptive information based on an observation of physical characteristics) or **quantitative** (numerical information). An example of a qualitative chemical property is when a pyrotechnic colorant like calcium chloride is added to a firecracker to make an orange flame. An example of a quantitative chemical property is how much time it takes for gas bubbles to form when baking soda and vinegar are mixed together (which may take six seconds).

Physical Properties

Physical properties are ones that we can observe without chemically changing a substance. For example, the hardness and colour of a substance are both physical properties. No new substances are formed when we make these observations. Remember that physical properties can also be either **qualitative** (such as the colour of gold being yellow or copper metal being shiny/lustrous) or **quantitative** (such as the density of H_2O being 1.0 g/cm^3 and the melting point of H_2O being 0° C).

There are three physical properties that we will study in more detail: density, compressibility, and diffusion.

Density

You probably have some knowledge about **density** and perhaps have even done some density calculations. We often try to demonstrate the idea of density by using comparisons such as this one:

Two identical boxes have the same volume but contain different materials. Box A contains lead (think of the heavy apron the dentist uses before taking your x-rays) while Box B contains feathers. You could predict that the box of lead would have a greater mass. Having a greater mass for the same volume tells us that lead is a denser material than feathers. Sometimes, simple language such as “it’s heavy for its size, or, it is very light for its size” provides a good way of thinking about high and low density.

It is important to note that every substance has its own density. Since we can use this characteristic to help us identify a substance, it is therefore a property. For example, aluminum (Al) has a density of 2.7 g/cm^3 (that is, 2.7 grams of mass for one cubic centimetre of volume, which is about the size of a sugar cube). If density is a property of aluminum, this means that every sample of pure aluminum should have the same density. Likewise, if we have a sample of an unknown metal and calculate its density to be 2.7 g/cm^3, we could determine that the metal is probably aluminum. This assumes, of course, that no other metals have a density close to 2.7 g/cm^3.

Compressibility

The term **compressibility** contains a couple of words that you probably recognize. The words “compress” and “press” might give you an indication of squeezing. Reducing the space between particles allows us to fit more particles in the same space. Some materials compress easily, like most gases, if they have lots of space between their particles. Solids and liquids have very little space between their particles and are therefore more difficult to compress.

Diffusion

Have you ever noticed that you can smell the lunch special from the cafeteria even though you are on the second floor of the school? Thanks to diffusion, smells can spread easily through the air. What is diffusion?

Diffusion is the movement of one substance through another. Diffusion is not limited to gases though. It also occurs commonly in liquids and sometimes in solids. That is why you can be lazy and not stir the drink crystals into your glass of water. Come back later and diffusion will have helped the colour and flavour of the drink crystals move through the water.

One rule to remember when it comes to diffusion is that substances diffuse from areas of high concentration to areas of low concentration. Also, lighter particles diffuse more quickly than heavier particles.

Now that we have refreshed our memories, let’s describe the **states of matter** and their physical properties in more detail.

Characteristics of the 3 Common States of Matter

Characteristics			
	Solid	Liquid	Gas
Shape	definite shape	takes the shape of its container	takes the shape of its container
Volume	definite volume	definite volume	takes the volume of its container
Density	usually very dense	usually less dense than solids	usually much less dense than solids and liquids
Compressibility	not easily compressed	not easily compressed	easily compressed
Diffusion	does not easily diffuse	easily diffuses	diffuses very easily

The Fourth State of Matter

A *fourth* state of matter? That's right! You are most likely familiar with the three most common states of matter (solid, liquid, and gas), but there is also one uncommon state of matter. This fourth state of matter is called **plasma**, and while some of you may know the word you may not know much about its characteristic properties.

What is Plasma?

Plasma can be defined as a gaseous mixture of positive ions and electrons. Perhaps you could imagine this as a large number of atoms of a substance having had some of their electrons stripped away. Plasma temperatures and densities range from relatively cool to very dense and hot. On Earth, this special mixture can only be created at very high temperatures in a laboratory (over 100 million degrees Celsius) and then carefully contained for use in objects like plasma TVs. Other types of plasmas, such as the aurora borealis (Northern Lights) at –170 degrees Celsius, can be quite cool.

The universe is made up of 99% plasma, but there is very little to be had here on Earth. For this reason, we call plasma an uncommon state of matter. You and I could observe some of the universe's plasma by watching lightning bolts and the stars in the night sky (whose plasmas are heated to very high temperatures by nuclear reactions inside the stars). Ordinary solids, liquids, and gases are generally too cool or dense to be in a plasma state.

To help you review what you have just learned about the four states of matter and their properties, you will now complete Learning Activity 1.1. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Learning activities are not to be sent in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 1.1

Characteristics of Matter

1. For each statement, determine if it is true or false. Try to correct all statements you think are false.
   a) Liquids are easier to compress than solids.
   b) Solubility is a chemical property.
   c) Flammability is a chemical property.
   d) Density is a ratio that compares the mass of an object to its volume.
   e) The density of a substance changes as the substance changes state.
   f) A chemical property does not involve a substance combining with or changing into other substances.
   g) Granulated sugar is a solid form of matter with a definite shape but no definite volume.
2. Can you think of any specific examples where substances are compressed?
3. Using your knowledge of compression, why does a tank of compressed oxygen allow scuba divers to stay underwater for long periods of time?

Right after you read the lesson summary, you will complete Assignment 1.1. It will help you apply your knowledge of the states of matter and their properties. This assignment (along with all other assignments) is worth marks. Unlike all other assignments in this module, you will submit it to the Distance Learning Unit immediately after you have finished rather than at the end of the module.

Lesson Summary

In this lesson, you learned about the four states of matter, including concepts such as density, compressibility and diffusion. In the next lesson, you will learn about the Kinetic Molecular Theory (which was introduced to you way back in Grade 7) and how it explains the properties of gases.

Assignment 1.1

Substance Investigation *(17 marks)*

In Part A of this assignment, you are asked to choose a substance and further investigate its properties. Choose any material at home or school that you can safely work with. Before you begin, you must email or telephone your tutor/marker and obtain feedback on your choice. Not only will the tutor/marker help you decide if the substance is a good choice for this assignment, but he/she can also suggest other options if necessary. This assignment is worth 17 marks.

Name: ______________________________

Part A

1. Choose a substance you are interested in investigating. Write the name of the substance that you have chosen on the line below. Now call or email your tutor/marker, and discuss your choice before going any further. *(1 mark for contacting your tutor/marker)*

   ____

2. a) What is the state of this substance at room temperature? *(1 mark)*

   ____

   b) For each of the following aspects, use a few words to qualitatively describe your substance more specifically and provide a few words to support your statement. *(1 mark* x *5 = 5 marks)*

   Shape:

   ____

   Volume:

   ____

   Density:

   ____

continued

Assignment 1.1: Substance Investigation (continued)

Compressibility:

Diffusion:

Part B

You have just been introduced to plasma—the fourth state of matter. Your job is to use the Internet to briefly research plasma and respond to the following questions using your findings. If you do not have access to the Internet, consult your school or community librarian for guidance as to how to complete your assignment. You may also contact your tutor/marker for additional help. *(10 marks)*

Before starting your research, take a look at the example provided below. This will give you an idea of the depth of detail you are expected to provide in your answers.

Example: Energy is required to strip electrons from atoms to make plasma. What happens to plasma that does not get enough sustaining power? *(1 mark)*

Answer: In this scenario, plasma particles will recombine and form neutral gases.

1. Describe the way that plasmas commonly look and explain why. *(2 marks)*

2. Compare the conductivity of plasma to that of metals regularly used in electrical wiring, like copper. *(1 mark)*

continued

Assignment 1.1: Substance Investigation (continued)

3. Give three examples for the uses of plasma and/or places where you would find naturally occurring plasma. *(3 marks)*

____

____

4. Is plasma a long-lasting state? Comment on the lifetime of plasma. *(1 mark)*

____

____

5. Comment on plasma's ability as a producer of electromagnetic radiation, compared to the other states of matter. *(1 mark)*

____

____

6. In this lesson, plasma was referred to as an "uncommon state of matter." Why, then, is plasma also referred to as "the most common form of matter known?" *(1 mark)*

____

____

7. Plasmas are synonymous with chaos and instability, but yet they are used in televisions and lighting. How does this happen? *(1 mark)*

____

____

NOTES

Submitting Your Assignments

It is now time for you to submit the Assignment 1.1 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 1 assignments and organize your material in the following order:

- Module 1 Cover Sheet 1 (found at the end of the course Introduction)
- Assignment 1.1: Substance Investigation

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

NOTES

Lesson 2: Kinetic Molecular Theory
(2.5 hours)

Lesson Focus

SLO C11-1-02: Use the Kinetic Molecular Theory to explain properties of gases.
Include: random motion, intermolecular forces, elastic collisions, average kinetic energy, temperature

SLO C11-1-03: Explain the properties of liquids and solids using the Kinetic Molecular Theory.

Lesson Introduction

In the previous lesson, you learned to describe the states of matter according to their physical properties. In this lesson, you will take that one step further by using the Kinetic Molecular Theory to explain the properties of gases, liquids, and solids.

Simply put, gases, liquids, and solids do not share all the same properties. For example, why is it almost impossible to compress a solid? Why do liquids flow more slowly when they are cooled? Understanding the Kinetic Molecular Theory will help you answer these, and many more, questions about the behaviours of gases, liquids, and solids.

What is the Kinetic Molecular Theory?

Chemists use the **Kinetic Molecular Theory** to explain why matter, especially gases, behaves as it does. The word "kinetic" means "to move," and objects in motion have energy called **kinetic energy**. Knowing this might help you see how the Kinetic Molecular Theory was named, since it describes the size, motion, and energy of particles. You might also recognize some of these terms from a Physics course. The major points of the theory are:

- The particles that make up matter are very small.
- There are spaces between the particles. Gas particles are much smaller than the distances between them. This means that almost all the volume of a gas is empty space, and explains why gases are so easily compressed.

- Particles of matter are in constant random motion. In gases, the particles are in constant straight-line motion. They only change direction when they collide with another particle or the sides of their container. Such collisions are completely elastic, meaning there is no loss of energy.
- There are forces of attraction between the particles in matter called **intermolecular forces**. In gases, these forces are negligible and gas particles neither attract nor repel each other.
- As temperature increases, the speed of the particles increases, and as temperature decreases, the particles' speed decreases. In other words, the kinetic energy of the particles increases with increasing temperature and decreases with decreasing temperature.

The points of this theory will make more sense as we examine gases and their behaviour.

Gases

Using the Kinetic Molecular Theory and our knowledge of physical properties from Lesson 1, we can develop a particulate model for gases.

Gases are easily compressed, as is demonstrated by air compressors, bicycle pumps, and aerosol cans. This means the particles of a gas must be very far apart to start out with. Gases also have low densities, suggesting the particles are quite loosely packed. In fact, in a container of gas, less than one-tenth of one percent of its total volume is occupied by gas particles. This means more than 99.9% of a container of gas is empty space!

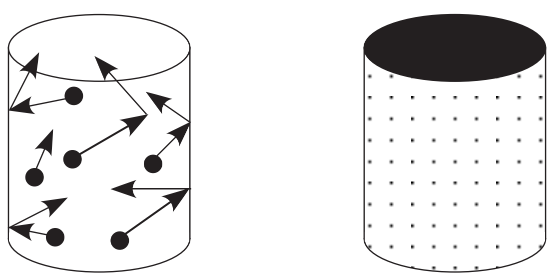


Figure A Figure B

Gases will fill any container in which they are placed and they will diffuse very easily. For example, a 200 mL sample of gas will spread out to occupy a 500 mL volume if it is moved into a 500 mL container.

Given the large spaces between gas particles (See Figure B above), there is room for particles of other gases to move past one another and mix together. Since gases will spread out and fill any container, the forces of attraction between the particles must be very low. If the forces of attraction were large, the particles would be held in a defined space, like solids and liquids. This means that the particles of a gas move freely in rapid straight-line motion and their motion is rarely changed by the interference of other gas particles. The freedom of motion also suggests that gas particles do not attract or repel each other strongly, if at all.

Let's examine in more detail how the Kinetic Molecular Theory explains the properties of gases:

Random Motion

According to the Kinetic Molecular Theory, gases are in constant random motion. This randomness allows gas particles to spread out and mix with each other. You might recall in Lesson 1 learning that diffusion can take place without physically mixing substances together. The random motion of gas particles causes two or more gases to mix until they are evenly distributed. The same random motion allows gas particles to spread out and fill the entire volume of the container in which they are held.

Intermolecular Forces

Intermolecular forces hold two separate particles together, such as two water molecules in a drop of water. Strong intermolecular forces will not allow particles to freely move. Since gas particles move easily and freely, we assume that they do not attract or repel each other. Intermolecular forces are weaker between particles that have more space between them.

Solids and liquids do not move as freely, and generally have a defined shape and volume, which indicates stronger intermolecular forces between their particles. As the particles of liquids and solids are closer together, the intermolecular forces between their particles are higher.

Elastic Collisions

According to the Kinetic Molecular Theory, the particles of a gas are constantly moving in random straight-line motion. If the gas particles are in a container, the particles must eventually collide with the sides of the container or other gas particles (See Figure A above). When the particles collide with the sides of the container, they exert a force upon the container's walls. We call this force, which acts on each square centimetre of area of the container wall, **gas pressure**. Pressure is defined as force per unit area.

There are three possible outcomes from these collisions between gas particles and the sides of their container:

1. A collision could result in a *loss of energy*—that is, the energy the particle contains before the collision (initial energy, or E_1) is greater than after the collision (E_2). If this were true and $E_1 > E_2$, the particles would eventually slow down due to energy loss and the pressure in the container would decrease.

2. If the particles have more energy after the collision than before, the particles would *gain energy* due to collisions. In this case, $E_1 < E_2$, and the gas particles would gain energy and speed up over time. The pressure in the container would increase in this situation.

3. If the energy of the particles before and after the collisions are equal ($E_1 = E_2$), the pressure inside the container would *remain constant*, but only provided the temperature does not change.

Let us consider these three scenarios by using a propane barbeque tank as an example. When you are finished cooking your hotdogs, you turn off the valve for the propane tank. Even though the propane tank is not used between barbeque sessions, the tank maintains its pressure (the force of the gas on the sides of the tank), provided the amount of gas remains constant (that is, there are no leaks) and the temperature remains unchanged. If the particles gained energy with every collision, the force of each collision would increase the pressure, which would be a potentially explosive situation. If the particles lost energy with every collision, the force of each collision would eventually decrease, resulting in a lower pressure and a tank that may not work the next time you wanted to use it. This is why you might see a warning label on a tank of compressed gas recommending a stable storage temperature.

According to the Kinetic Molecular Theory, collisions between particles and between particles and their container are perfectly **elastic**. This means there is no loss of energy. This is illustrated in the diagrams below.

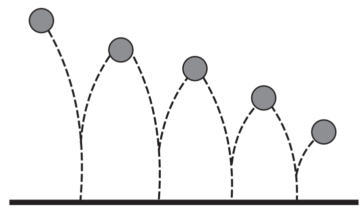


Figure 1: Inelastic Collisions

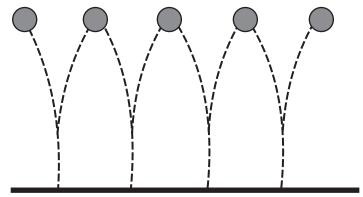


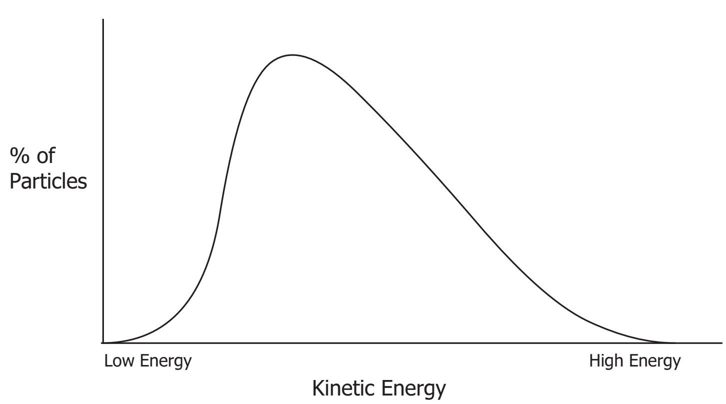
Figure 2: Elastic Collisions

In Figure 1, the bouncing ball loses energy with each bounce, and as a result the force of the bounce decreases as does the height of its bounce. In Figure 2, the bounces are perfectly elastic. There is no loss of energy, so the ball always returns to its original height. This illustrates the movement of gas particles.

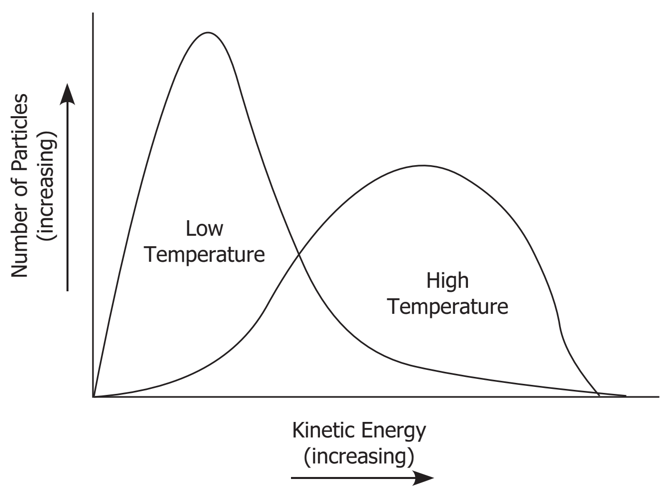
Average Kinetic Energy

In 1866, Clerk Maxwell and Ludwig Boltzman performed experiments to study the kinetic energy of gaseous substances. They found that not all particles, at a given temperature, had the same amount of energy. In fact, they found that some had very little energy, some had a lot, and most were somewhere in between. Since most of the particles have kinetic energy somewhere in the middle of the range, this is considered to be the **average kinetic energy**. A diagram illustrating this, such as the one below, is called a Kinetic Energy Distribution Curve. You can see that some particles have either very high or very low values of kinetic energy, but the majority of the particles have an intermediate amount of kinetic energy. This average kinetic energy is represented on the following graph by the peak.

Kinetic Energy Distribution



The graph below shows the relationship between temperature and kinetic energy. The first curve shows the distribution of kinetic energies for particles at a lower temperature. The second curve shows the energy distribution for the *same* group of particles at a higher temperature. In both cases, most of the particles have intermediate kinetic energies, which are close to the average value. However, there is a wider range of kinetic energies at higher temperatures, which is why the second curve is wider with a lower peak. What does this type of graph really tell us about kinetic energy? It says that an increase in the average kinetic energy of a sample of particles causes the temperature of the substance to increase, which results in a curve that flattens toward the right side of the graph. At lower temperatures, kinetic energy decreases, resulting in a taller curve found closer to the left-hand side of the graph.



Each particle in a sample of matter has a different amount of kinetic energy, so they will all move at a different speed. Kinetic energy is calculated using the expression $\frac{1}{2}mv^2$, where "v" is the velocity (or speed) of each particle and "m" is the mass of the particle. Since velocity is squared, changing the temperature changes the speed of each particle by a different ratio and the curve changes shape.

Liquids

In Lesson 1, you revisited the phases of matter. You might remember that liquids take the shape of the container in which they are placed and have a fixed volume. In other words, the particles in a liquid can spread out and adjust to the shape of the vessel, but the space between the particles does not change. If the space between particles stays the same, then the volume of a liquid also stays the same, unlike gases. For example, 200 mL of milk in your glass can spread out to make a big puddle if you spill it. However, it can only be a 200 mL puddle of milk!

Like gases, liquid particles are also in constant motion. Their motion, however, is limited, since the intermolecular forces between them are stronger than their gaseous counterparts. It is also for this reason that liquids are not easily compressed and have relatively high densities.

Model of Liquids

A model of the particles in a liquid may look like the illustration below:

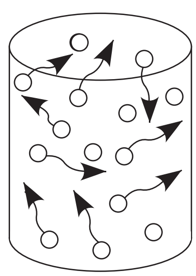


Figure A

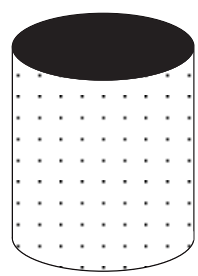


Figure B

Note that the particles in a liquid are closer together than those found in gases, but are still further apart than those in a solid (See Figure B). Less space and medium kinetic energies allow these particles to flow past one another (See Figure A). Thus, liquids, like gases, are fluid substances. You might be surprised to learn that liquids flow less easily than gases, though. This is because of the stronger intermolecular forces found between liquid particles. You have likely encountered these intermolecular forces at work when trying to pour a thick liquid, like ketchup or syrup. These liquids do not flow as quickly as milk or water. Why is there such a difference? **Viscosity** is the measure of resistance of a liquid to flow. The more resistance there is, the slower the flow of the liquid. This resistance comes from the intermolecular forces between the particles.

Viscosity is affected by size and shape of particles, as well as temperature. At colder temperatures, the kinetic energy of particles decreases and the viscosity of that liquid increases. This is why engine oil does not flow well on a cold day and needs to be heated by a block heater. With higher temperatures, viscosity decreases as kinetic energy increases, which makes particles in motion able to overcome their intermolecular forces. This is why molasses will flow more quickly when heated for cooking or baking.

Solids

Since solids are very difficult to compress, their particles must be so close together that they cannot be easily forced any closer. The high density of most solids also suggests there are more particles per unit volume than both liquids and gases.

The definite shape and volume of solids also implies that the particles are held together quite tightly and must not move very much. In fact, they probably just vibrate in place. Therefore, the forces of attraction between particles (or **intermolecular forces**) in a solid must be very high. It should be noted that forces of attraction and chemical bonds are not the same.

Model of Solids

A model of the particles in a solid may look like the illustration below:

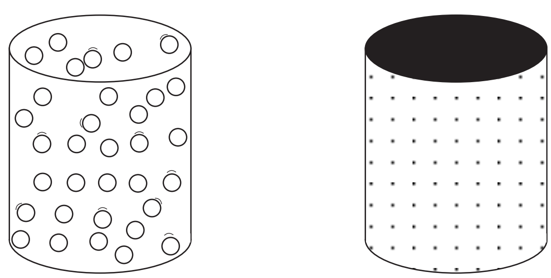


Figure A   Figure B

Note that the particles in a solid are much closer together (See Figure B) than those found in liquids and gases. A lack of space and low kinetic energies allow these particles to simply vibrate (See Figure A).

In order for diffusion to occur, particles must easily move through the matter. Since the particles in a solid are very close together, other particles cannot easily move through and mingle with solid particles. While solids do not generally diffuse through solids or liquids, they can sublimate and then diffuse through other gases. This is what happens when you use a solid air freshener in one room, but the scent travels through the house to other rooms.

If you have access to a computer, the animation found at www.harcourtschool.com/activity/states_of_matter/ summarizes the differences in spacing and movement of the particles in the solid, liquid, and gaseous states.

To help you review, or practise, what you have just learned about the Kinetic Molecular Theory, you will now complete Learning Activity 1.2. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Learning activities are not to be sent in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 1.2

The Kinetic Molecular Theory

1. Complete each of the following statements by finding the important missing term(s).
    a) The Kinetic Molecular Theory helps us understand how the ______ , ______ , and ______ of particles explain the behaviour of matter.
    b) Particles that move have __________ .
    c) ______ particles tend to have the most kinetic energy, since they have more space between them to move.
    d) Since gas particles are separated by empty space, gases have a high degree of __________ .
    e) ________ are also compressible, but not as easily as gases. Therefore, their degree of compressibility is less than gases.
    f) Since liquids are not easily compressed, their ________ tends to be fixed.
    g) ________ also have a fixed volume, since their particles cannot be forced any closer together than they already are.
    h) Solids can be described by some of these __________ : high density, low kinetic energy, and high intermolecular forces.
    i) The properties of all states of matter can be explained by the ________ .
2. The following is a popular science demonstration: A few drops of water are placed in an aluminum soft drink can. The can is then placed over a flame and heated to a high temperature for a minute or so. The aluminum can is then quickly turned upside down and submerged into cold water. The can implodes and crumples before your eyes. Use the Kinetic Molecular Theory to help explain why the can collapsed. If you have access to the Internet and would like to see this demonstration, you can check out the video at www.youtube.com/watch?v=1Efy36yxdUc. You should also be able to find it by using the search term "can crush" in the youtube search window.
3. Use your own words to describe the basic assumptions of the kinetic molecular theory of gases as far as: (a) volume, (b) intermolecular forces, and (c) collisions.

Right after you read the lesson summary, you will complete Assignment 1.2. It will help you review the unique properties of gases, liquids, and solids. This assignment (along with all other assignments) is worth marks. You will submit it, along with the rest of the other assignments in this module, to the Distance Learning Unit when you have finished the module.

Lesson Summary

In this lesson, you learned about the Kinetic Molecular Theory and how it explains the unique properties of gases, liquids, and solids. In the next lesson, you will learn how the same theory helps us explain phase changes.

NOTES

Assignment 1.2

Properties of Gases, Liquids, and Solids *(20 marks)*

1. Complete the following sentences using the appropriate term. *(1 mark each)*

   a) An ________________ is one in which no kinetic energy is lost.

   b) ________________ are the type of matter most easily compressed.

   c) The kinetic energy of particles usually ________________ as temperature increases.

   d) Gases can easily expand due to ________________ motion of their particles.

   e) Stronger intermolecular forces usually result in liquids having higher ________________.

   f) ________________ have a definite shape and volume because their particles can only vibrate around fixed points.

   g) Gases and liquids are both considered to be ________________.

2. Explain WHY both gases and liquids are considered to be fluids. Use the Molecular Kinetic Theory in your answer. *(2 marks)*

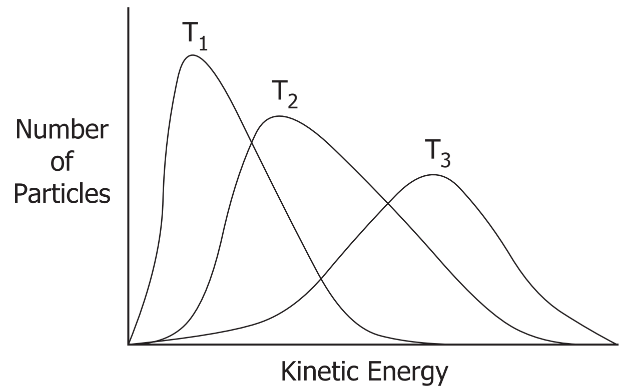
   ____

   ____

   ____

continued

3. Use the Kinetic Energy Distribution curves below to answer the questions.



a) Which point on each curve represents the average kinetic energy? *(1 mark)*

____

____

b) Use your own words to explain the shapes of the curves for each temperature curve. *(1 mark* x *3 curves* = *3 marks)*

____

____

c) What do you know about the kinetic energies of each of the three curves of this graph? *(1 mark each* = *3 marks)*

____

____

d) Predict what would happen to the position and shape of the next curve if the temperature was even higher than T3. *(2 marks)*

____

____

continued

e) In which sample is the average kinetic energy of the particles highest? *(1 mark)*

____

____

f) Why do particles at cold temperatures tend to have the lowest average kinetic energies? *(1 mark)*

____

____

Notes

Lesson 3: Phase Changes (2 hours)

Lesson Focus

SLO C11-1-04: Explain the process of melting, solidification, sublimation, and deposition in terms of the Kinetic Molecular Theory.
Include: freezing point, exothermic, endothermic

SLO C11-1-05: Use the Kinetic Molecular Theory to explain the processes of evaporation and condensation.
Include: intermolecular forces, random motion, volatility, dynamic equilibrium

Lesson Introduction

In Grade 7 Science, you learned about changes of state. What really happens when these changes occur? Can you use the Kinetic Molecular Theory to explain why, for example, boiling water results in steam production and less water in the kettle? In the previous lesson, you learned about the Kinetic Molecular Theory and how it helps us understand the behaviour of solids, liquids, and gases. In this lesson you will find out how particles behave during these changes, as well as the role kinetic energy plays in the process.

Phase Changes

In Grade 7 Science, you were introduced to the particle theory of matter. You might remember learning about how particles move as a liquid changes to a gas or a solid. At that time, you probably used the terms "phase" and "state." Are the words "phase" and "state" the same thing? Actually, they are not. Most substances can exist in three states of matter, depending on their temperature and pressure. For example, water can exist in the gaseous (vapour), liquid (water), and solid (ice) states. If, however, you have water and ice together in a glass, then we use the term "icewater" to describe the phase. A **phase** refers to the mixture of states of matter that coexist as physically distinct parts of a mixture.

You may remember that changes of state require that energy be *added* or *removed*. The form of energy that you probably talked about the most in Grade 7 was heat energy. This energy has an important job, and it is not simply to increase or decrease temperature. Rather, the heat energy is used to break and re-form the intermolecular forces that hold molecules together. Before we study phase changes in more detail, let's first learn about two important energy terms that will be used frequently in this lesson.

Endothermic phase changes are those that require energy to be added to the system for the change to take place. The energy required allows intermolecular forces to be overcome such that they can be broken and molecules can move further apart from each other. Pulling particles further apart requires work, and work requires energy. **Exothermic** phase changes are those in which more energy is released than is required to break bonds. Where does this energy come from? The energy is stored in particles as potential energy.

Keep all of this in mind as we move on to study the phase changes below.

Melting

A solid will change to a liquid when heat is added or absorbed. This change is called **melting** or **fusion**. Energy is needed to pull the particles apart and overcome the forces of attraction in the solid. This added energy remains stored in the particles that are now further apart. The amount of energy required to melt a solid depends on the strength of the forces keeping the particles together in the solid. As a solid is heated (energy is absorbed), the forces of attraction between the particles are slowly overcome, allowing some particles to move more freely. As more particles are free to move, the substance melts. Since energy is required to overcome the forces of attraction, this is an example of an endothermic reaction.

The **melting point** of a substance is the temperature at which a substance melts or fuses. It is unique for every substance.

Solidification

Freezing or **solidification** is the conversion of liquid to solid. In this case, heat is removed from the substance and the molecules lose kinetic energy. As the particles slow down and move closer together, potential energy is released. As a substance freezes (energy is lost), the particles continue slowing down and the forces of attraction between the particles, or intermolecular forces, begin to increase and take hold. As the intermolecular forces increase, the particles arrange themselves into an organized repeating pattern called a crystal. Since this type of phase change releases energy, it is an example of an exothermic process.

The **freezing point** of a substance is the temperature at which a substance freezes. It is unique for every substance. For most substances, the freezing point and the melting point are the same.

Sublimation

Sublimation occurs when a solid changes directly to a gas, without passing through the liquid state. Solid iodine, mothballs, and carbon dioxide (dry ice) all sublime at room temperature. Dry ice is particularly useful as it does not get "wet" with liquid. This is quite convenient for storing frozen foods, as the food containers do not get wet from melting ice. Snow disappears in the winter due to sublimation. Some winters, more snow disappears due to sublimation than melting! As you may know, before clothes dryers were popular, people would hang clothes outside to dry after washing. In the winter the clothes would freeze, then eventually dry by sublimation. Most freeze-dried foods are first frozen at very low temperatures and the water sublimes after being attached to a vacuum.

Since sublimation requires that energy be added, it is an example of an endothermic process.

Deposition

Deposition is the opposite of sublimation. It is the direct conversion to a solid from a vapour without forming a liquid. In the winter, it is too cold for the formation of a liquid, yet ice is formed on trees and your car windshield from water vapour in the air. Thus, the formation of frost in the winter, or in your freezer, is a good example of deposition. When water vapour high in the atmosphere changes directly into a crystalline solid, you will see snowflakes falling. Since energy is released as the solid forms, this is an example of an exothermic process.

Evaporation

Energy is needed to change a liquid to a gas. Since the particles of a gas are much farther apart than in a liquid, the intermolecular forces must be overcome to allow the particles to move more freely. Overcoming these forces of attraction requires the input of energy in order to increase the kinetic energy of the liquid molecules. This energy is stored in the liquid molecules as potential energy. Particles that obtain enough energy to escape from the liquid and enter the gas phase have participated in a change known as **vaporization**.

Evaporation is the conversion of a liquid to a gas that occurs specifically on the surface of a liquid. This process generally requires energy to be added gradually. You have likely witnessed evaporation when a glass of water left out on a table appears less full a few days later. This is the same phenomenon that allows your wet hair to dry in the air after a shower.

Substances that evaporate quickly are said to be **volatile**. Volatile substances, such as paint thinner and alcohol, have weak intermolecular forces of attraction that are easily overcome. These substances also disperse into the air quickly, such that you can smell them rapidly as they spread through the air. You might have observed this phenomenon when you or someone in your house has used nail polish remover. If left uncapped, the polish remover would evaporate faster than the same volume of water, and you would be able to smell the distinct scent quickly as the nail polish remover particles disperse throughout the house.

Since evaporation requires energy, it is an example of an endothermic process.

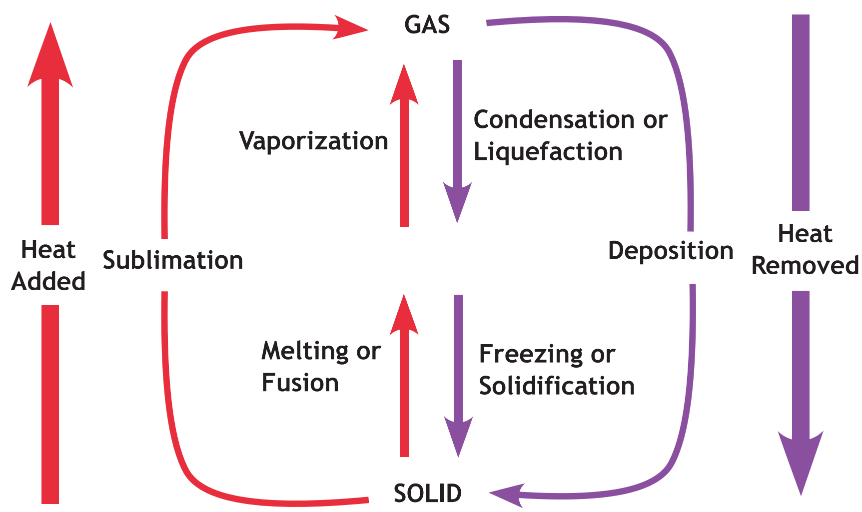
Condensation

The conversion of a gas to a liquid is called **condensation** or **liquefaction**. As a gas cools, the kinetic energy of the particles decreases. The particles slow down and their kinetic energy can no longer resist the forces of attraction from other particles. The particles are drawn closer together by the intermolecular forces, and the potential energy that was stored is released into the surroundings. Since condensation releases energy, it is an example of an exothermic process. An example of condensation is when a layer of air near the ground cools and water vapour condenses to create a fog.

In a closed container, an interesting phenomenon occurs. Think of a container of leftover soup that is still warm when you put it in the refrigerator. As the warm air in the container cools, condensation will collect on the sides and underneath the lid of the container. Evaporation will occur on the surface of the soup. If this container stayed closed and at a constant temperature, eventually for every particle moving from the liquid to the gaseous phase, there would be one molecule moving from the gaseous to the liquid phase. The rate of condensation and the rate of evaporation would then be equal. This is called **dynamic equilibrium**, where

Rate of evaporation = *rate of condensation*

Use the diagram below to help you organize all of the phase changes that we just reviewed.



To help you review, or practise, what you have just learned about phase changes, you will now complete Learning Activity 1.3. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Learning activities are not to be sent in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 1.3

Phase Changes

1. What phases are in equilibrium at a substance's melting point?
2. Compare the evaporation of a liquid in a closed container with that of a liquid in an open container. Use the term "dynamic equilibrium" in your answer.
3. Perspiration during exercise or on a hot day is an example of
   a) Condensation
   b) Sublimation
   c) Evaporation
   d) Deposition
4. Which phase changes are endothermic? Which are exothermic?
5. The vaporization of a solid is also known as
   a) Condensation
   b) Deposition
   c) Evaporation
   d) Sublimation
6. Water vapour in the atmosphere will _______ to form rain.
   a) Sublime
   b) Condense
   c) Evaporate
   d) Solidify

Right after you read the lesson summary, you will complete Assignment 1.3. It will help you organize important information about phase changes. This assignment (along with all other assignments) is worth marks. You will submit it, along with Assignments 1.2, 1.4, and 1.5, to the Distance Learning Unit when you have finished the module.

Lesson Summary

In this lesson, you learned how our knowledge of the Kinetic Molecular Theory can help us understand phase changes. In the next lesson, you will learn about vapour pressure and how it can be measured.

Notes

Assignment 1.3

Dynamic Equilibrium and Phase Changes *(10 marks)*

1. In your own words, explain dynamic equilibrium. *(1 mark)*
   What is happening to the molecules when this occurs? *(1 mark)*

2. Would you expect dynamic equilibrium to be achieved in an open container? Explain your answer. *(2 marks)*

3. When you remove the lid from an ice cream container that has been in the freezer for a couple of months, you notice ice crystals on the underside of the lid. The crystals form after there is: *(1 mark)*
   a) Sublimation
   b) Deposition
   c) Condensation
   d) Evaporation

continued

Assignment 1.3: Dynamic Equilibrium and Phase Changes (continued)

4. When a solid is heated to its melting point, which of the following would you **NOT** expect? *(1 mark)*
   a) Condensation will likely occur.
   b) Kinetic energy will increase.
   c) Some particles will be able to overcome the intermolecular forces.
   d) An endothermic reaction takes place.

5. Compare melting and condensation based on kinetic energy changes *(2 marks)* and spacing of particles. *(2 marks)*

   ____

   ____

   ____

   ____

Lesson 4: Vapour Pressure (2 hours)

Lesson Focus

SLO C11-1-06: Operationally define vapour pressure in terms of observable and measurable properties.

SLO C11-1-07: Operationally define normal boiling point temperature in terms of vapour pressure.

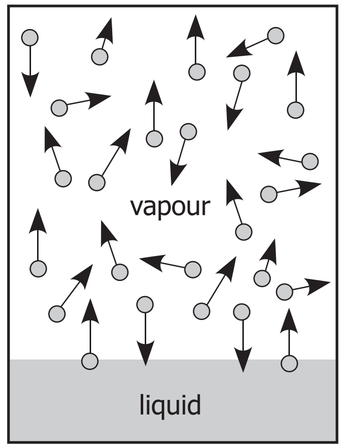
Lesson Introduction

An open container of water is placed next to a closed container of water. Let's assume that both containers start with the same volume of liquid. How will evaporation take place in both containers? Will the closed container of water eventually evaporate completely? How does temperature affect evaporation? In the previous lesson, you learned about phase changes and how they are related to the Kinetic Molecular Theory. In this lesson, you will focus specifically on one phase change, evaporation, and the variables that affect the pressure created by this process.

What is Vapour Pressure?

If a container of liquid is left open, the liquid will completely evaporate. This is because the rate of evaporation of the liquid will be greater than the rate of condensation, and so over time all of the liquid will convert to the gaseous state. However, the liquid in the container that is closed will not evaporate completely. Do you know why? Since the vapour cannot escape into the open air, it collects inside the container and exerts pressure on the remaining liquid. Eventually, the space above the liquid becomes saturated with vapour and for every molecule that evaporates, another must condense.

The diagram below illustrates what would happen in a closed container:



Liquid in a closed container, at a constant temperature, reaches dynamic equilibrium with its vapour. This means that the liquid escapes and re-enters the liquid state at the same rate (this was shown in the diagram). At dynamic equilibrium, we said that

Rate of evaporation = rate of condensation

The pressure created by the vapour at equilibrium is known as the **vapour pressure**, abbreviated P_{vap}. In other words, vapour pressure is the pressure exerted by a vapour over a liquid. Vapour pressure is a physical property.

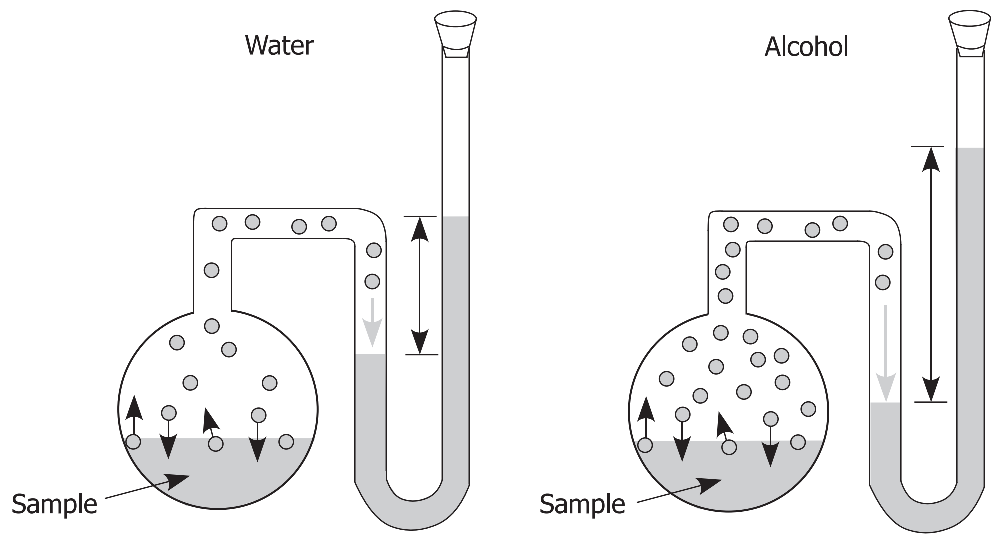
Vapour Pressure and Intermolecular Forces

The strength of intermolecular forces between the particles of a sample determines the rate of evaporation. As different substances are held together by different intermolecular force strengths, the rates of evaporation of different substances will be different. Try putting a few drops of rubbing alcohol and water on the countertop, for example. Which evaporates faster? The alcohol. Since the temperature of both liquids is the same (room temperature), their intermolecular forces must determine which evaporates faster. If the alcohol evaporates faster, it must then have weaker intermolecular forces than the water.

If intermolecular force strength is weaker, the amount of vapour that is produced increases. It is therefore easier for the particles of liquid to escape the forces of attraction in the liquid and become gaseous.

Measuring Pressure

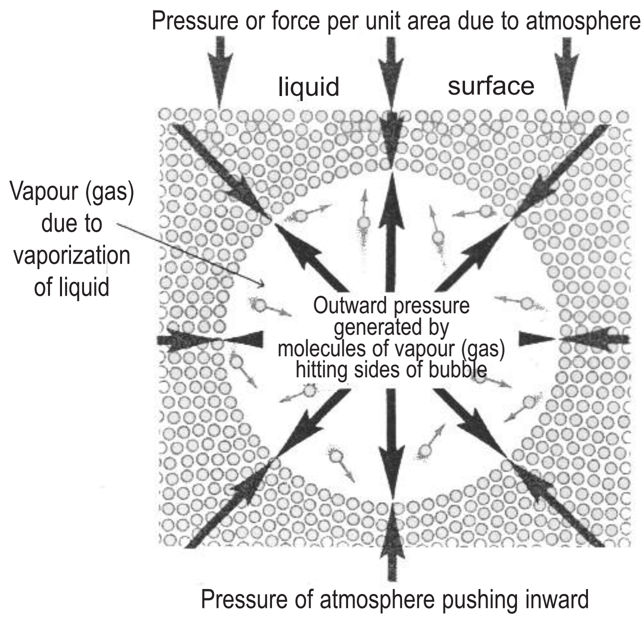
Pressure can be measured using several devices or instruments. A **manometer** usually has a bulb or glass container on one end and can be open or closed on the other. A liquid, often mercury, is placed in its U-shaped tube. The pressure of a sample in the bulb can be measured by finding the difference in height on both sides of the tube. As pressure increases in the bulb, the mercury is forced higher up the U-shaped tube. In the diagrams below, both U-tubes are at the same temperature.



Notice that the manometers above show the difference in the height of the mercury to be much greater with alcohol in the vessel, as compared to water in the vessel. Generally, the greater the forces of attraction, the lower the vapour pressure. This is why the water, having greater forces of attraction, shows a smaller difference in liquid levels in the U-shaped tube and thus has a lower vapour pressure.

Vapour Pressure and Boiling Point

In an earlier lesson you learned about vaporization. There are two types of vaporization, *evaporation* and *boiling*. Evaporation is the conversion of a liquid to gas on the surface of the liquid. **Boiling** occurs when vaporization takes place throughout the liquid and the *temperature of the liquid remains constant*. Bubbles of vapour collect below the surface of the liquid and then rise to the surface. This is known as an *operational definition* of boiling. The **boiling point** of a substance is defined as the temperature at which the *vapour pressure of a substance is equal to the external or atmospheric pressure*.



Atmospheric Pressure and Boiling Point

Since boiling occurs when the atmospheric pressure and the vapour pressure of a given liquid are equal, any change in atmospheric pressure will produce a change in the boiling point. If this is the case, then as atmospheric pressure decreases, so will the boiling point. Since less vapour pressure is required to match a lower atmospheric pressure, less heat is needed by the system.

This explains why it requires longer cooking time to prepare foods high above sea level. At higher altitudes, the atmospheric pressure is lower because the weight of the column of air above that area is less. The lower the air pressure, the lower the boiling point. So, if we were to try to hard boil an egg at a high altitude, the water would boil at a much lower temperature. Since the water would be boiling at that lower temperature, the egg would take much longer to cook because more heat would have to be delivered to it before the egg could become hard-boiled.

The **normal boiling point** of a substance is defined as the temperature when *vapour pressure equals standard pressure* (1 atmosphere, which is the same as 101.3 kPa, 760 mmHg, and 760 torr). See the table below for examples of normal boiling points at different altitudes.

Altitude	Normal Boiling Point of Water	Geographic Example
Sea Level	100° C	Key West, United States (1 m)
610 metres	98° C	Edmonton, Canada (668 m)
1524 metres	95° C	Guatemala City, Guatemala (1530 m)
2286 metres	92° C	Val Thorens, France (2300 m)
3048 metres	90° C	Haleakala, United States (3055 m)

Pressure cookers are often used to speed up cooking. A pressure cooker operates by increasing the pressure inside the pot. As pressure increases, the boiling point of the water also increases because more energy is needed to create more vapour so that the vapour pressure can become equal to the atmospheric pressure. The higher the temperature at which water boils, the shorter the cooking time, because a greater amount of heat will have already been delivered to the food by the time boiling occurs.

To help you review what you have just learned about vapour pressure, you will now complete Learning Activity 1.4. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Learning activities are not to be sent in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 1.4

Vapour Pressure

1. Which of the following statements does not describe what occurs as a liquid boils?
   a) The temperature of the liquid increases.
   b) Energy is absorbed by the particles.
   c) The vapour pressure of the liquid is equal to atmospheric pressure.
   d) Liquid particles are entering the gaseous phase.
2. Complete the following statement: Water boils below 100^{o} C on the top of a mountain because ...
3. In order for a liquid to boil, particles throughout the liquid must have enough ____________________ to vaporize.
4. What would you expect the effect that increasing temperature has on vapour pressure?
5. Why is the equilibrium that exists between a liquid and its vapour in a closed container called dynamic equilibrium?
6. There is a liquid-vapour equilibrium in a sealed container. If the volume of the liquid is increased, and the container re-sealed, how will the vapour pressure be affected?
   a) The vapour pressure will increase.
   b) The vapour pressure will decrease.
   c) The vapour pressure will initially increase, and then decrease.
   d) The vapour pressure will not change.

Right after you read the lesson summary, you will complete Assignment 1.4. It will help you review key points about vapour pressure. This assignment (along with all other assignments) is worth marks. You will submit it, along with Assignments 1.2, 1.3, and 1.5 to the Distance Learning Unit when you have finished the module.

Lesson Summary

In this lesson, you learned about vapour pressure and the correlation between vapour pressure and boiling point. In the next lesson, you will apply this knowledge to help you analyze pressure versus temperature graphs.

NOTES

Assignment 1.4

Vapour Pressure Problems *(10 marks)*

1. Explain why the boiling point of a liquid varies with atmospheric pressure. *(1 mark)*

2. A liquid is in a closed container and has a constant vapour pressure. What is the relationship between the rate of evaporation of the liquid and the rate of condensation of the vapour in the container? *(1 mark)*

   a) The rate of evaporation is faster than the rate of condensation.

   b) The rate of evaporation is slower than the rate of condensation.

   c) The rate of evaporation and the rate of condensation are equal.

3. Describe how the following three terms are related: evaporation, vapour pressure, and boiling point. *(3 marks)*

continued

Assignment 1.4: Vapour Pressure Problems (continued)

4. As intermolecular forces strengthen, would you expect the vapour pressure of a liquid to increase or decrease? *(1 mark)*

   Explain your answer. *(1 mark)*

5. Why are pressure cookers recommended for cooking at high altitudes? *(1 mark)*

6. Two sealed containers contain water and are at the same temperature. One jar contains 100 mL of water, while the other contains 15 mL of water. Despite the difference in volume, the vapour pressure in both containers is the same. Why is this? *(2 marks)*

Lesson 5: Graphing Vapour Pressure
(2 hours)

Lesson Focus

SLO C11-1-08B: Interpolate and extrapolate the vapour pressure and boiling temperature of various substances from pressure versus temperature graphs.

Lesson Introduction

Do all substances boil at 100º C, like water? No, liquids don't always boil at the same temperature, because a liquid boils when its vapour pressure is equal to the atmospheric pressure. This lesson will give you the opportunity to use your knowledge of vapour pressure and boiling point. After completing this lesson, you should be able to predict the vapour pressure and boiling temperature for a given substance by analyzing its vapour pressure curve.

Vapour Pressure: Important Points to Remember

We have already determined that intermolecular forces affect boiling point. We also determined that a liquid has reached its boiling point when its vapour pressure is equal to its atmospheric pressure. The statements below summarize some key ideas from the last lesson:

1. Intermolecular forces influence vapour pressure. Greater intermolecular forces result in lower vapour pressures.
2. A lower vapour pressure means that a greater amount of energy is required to create vapour over a liquid. Likewise, a higher vapour pressure indicates that less energy is required to create vapour.
3. The energy required to create vapour influences the boiling point of a liquid. As the energy requirements increase, the boiling point of a liquid increases.
4. Boiling point is also affected by intermolecular forces. A higher boiling point means that there are greater intermolecular forces that the liquid particles must overcome to vaporize.

Now you will have the opportunity to apply some of your knowledge by interpreting *vapour pressure versus temperature* graphs. Before starting the graphing portion of this lesson, some review is probably necessary to help you interpret the graphs you will be working with.

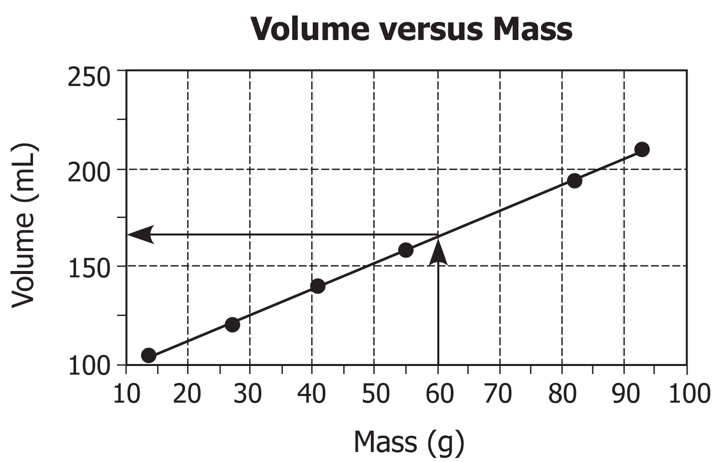
Interpolation and Extrapolation

Graphs are very useful tools, since we can use the curve or line to make further predictions regarding the relationship between variables. These techniques are not unique to chemistry, but are the same ones you use for any graph in any class. Let's quickly review two important graphing techniques that you will need to use to complete this lesson.

Interpolation

Sometimes we might need to look at the graph and use the line or curve to find information other than the points that are plotted. **Interpolation** is a technique that allows us to predict values of the independent and dependant variables. Using the line of best fit, we can determine a certain value even if it falls *between* measured points. This method is fairly reliable.

A quick example will help you remember what this really means. Take a look at Graph 1 below.



There are six measured points that are easy to read on Graph 1. However, sometimes you will need information from between these nicely plotted points. For example, what if you needed to know the volume that corresponds to 60 g of this substance? The first step is to draw a straight vertical line connecting the 60 g value on the *x*-axis to the line of best fit. If you then draw another straight line from this point, and connect it back to the *y*-axis, you can read the volume to be approximately 162 mL.

To help you practise what you have just learned about interpolation, you will now complete Learning Activity 1.6. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Learning activities are not to be sent in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 1.5

Interpolation

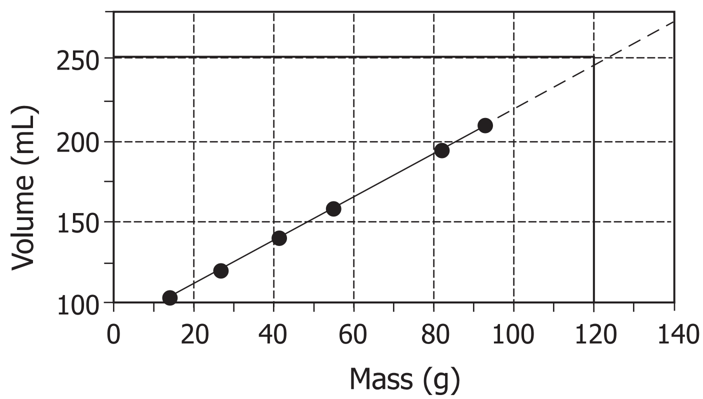
Using Graph 1 above, interpolate to get the following values:

1. The mass of 150 mL of this substance
2. The volume of 70 g of this substance
3. The volume of 30 g of this substance
4. The mass of 200 mL of this substance
5. The volume of 20 g of this substance

Extrapolation

In other instances, we may need to estimate a trend beyond the line or curve of the graph. **Extrapolation** involves using the line of best fit and extending it, using a broken line. By extending the line of best fit we can predict data beyond measured points from the data table. This method works best when the curve is not likely to change. Look at Graph 2 below:

Volume versus Mass



This is the same graph as you used in the previous example, with the same six points plotted. You can see that the line of best fit has been extended past the last point of the line of best fit. By doing this, it is possible to predict values past the measured values, assuming the trend of the graph does not change.

For example, let's predict the volume of 120 g of this substance. Again, the first step is to draw a straight, vertical line connecting 120 g on the x-axis with the line of best fit. Then connect this point with the y-axis, using a straight, horizontal line. The value for volume is about 250 mL.

To help you practise what you have just learned about extrapolation, you will now complete Learning Activity 1.6. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. And, like all other learning activities, you will not be sending it in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Keys found at the end of this module.

Learning Activity 1.6

Extrapolation

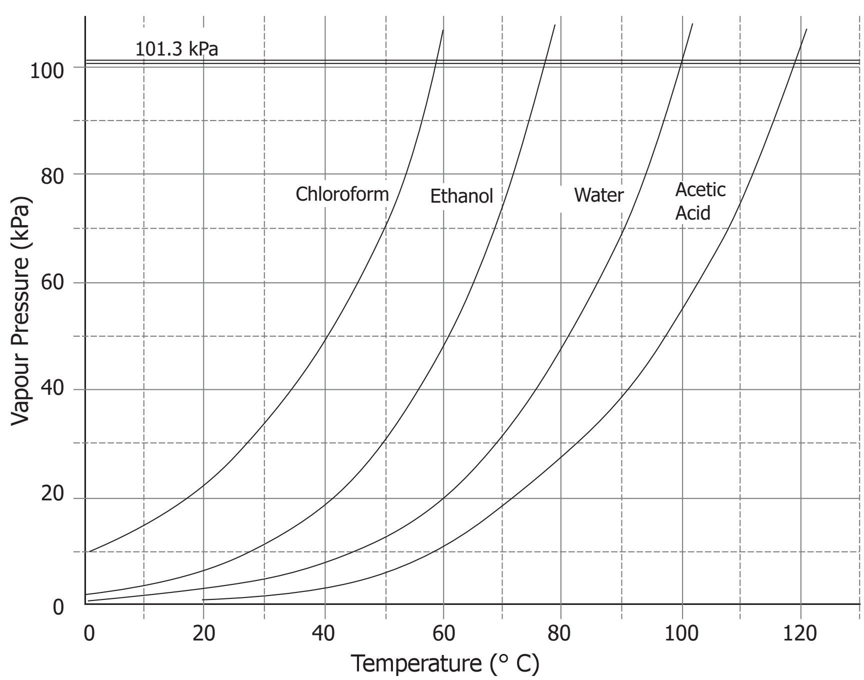
Using Graph 2 above, extrapolate to get the following values:

1. The volume of 110 g of this substance
2. The volume of 100 g of this substance
3. The mass of 275 mL of this substance

Vapour Pressure Curves

We can plot the vapour pressure of different substances as a function of temperature and get a graph like the one below. We will refer to this as Graph 3. Keep in mind that 101.3 kPa is standard atmospheric pressure.

Increasing Vapour Pressure with Temperature of Four Liquids



You will notice that for all four substances there is a similar trend—vapour pressure increases as temperature increases.

We can use Graph 3 to determine the normal boiling point of each substance and the temperature at which each will boil when pressure varies.

Example 1
What is the normal boiling point of ethanol?

Solution
From the graph, we look for the temperature at which the vapour pressure of ethanol is 101.3 kPa (standard pressure). This is using the interpolation method. *According to the graph, the normal boiling point is about 77° C.*

Example 2
If the pressure in a city is 90 kPa, at what temperature will water boil?
Hint: Use interpolation.

Solution
When the atmospheric pressure is 90 kPa, water will boil when its vapour pressure is 90 kPa. Start at 90 kPa on the y-axis and move to where it intersects with the curve for water, then see what the corresponding temperature is. *According to the graph, water boils at about 97° C at 90 kPa of pressure.*

Example 3
Which substance has the largest forces of attraction at 60° C?

Solution
The substance with the largest forces of attraction will have the lowest vapour pressure. *The substance with lowest vapour pressure at 60° C is acetic acid.*

Example 4
Which substance has particles that have the greatest intermolecular strength?

Solution
Acetic acid has the highest boiling point, which means that the liquid particles must overcome high intermolecular forces to vaporize.

Example 5
What is the boiling point of chloroform at 101.3 kPa?

Solution
Using interpolation, you should have found the boiling point to be around 58° C.

Example 6
What is the vapour pressure of water at 40° C?

Solution
Using interpolation, you should have found the vapour pressure to be about 8 kPa.

Example 7
Predict the vapour pressure of water at 110° C.

Solution
Using extrapolation, you might predict the vapour pressure to be about 125 kPa.

To help you practise what you have just learned about vapour pressure curves, you will complete Learning Activity 1.7. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. And, like all other learning activities, you will not be sending it in to the Distance Learning Unit for assessment. Check your answers against those provided in the Learning Activity Keys found at the end of this module.

Learning Activity 1.7

Interpreting a Vapour Pressure Graph

Use the vapour pressure graph (Graph 3) within the lesson to answer the following questions.

1. What would be the boiling point of water on a day when the atmospheric pressure is 95 kPa?
2. Ethanol is heated in a container in which there is a partial vacuum. The air pressure in the container is 25 kPa. At what temperature will the ethanol boil?
3. If substance “X” had a normal boiling point of 30° C, where would you expect to find the vapour pressure curve of “X”?
   a) To the left of the chloroform curve.
   b) Between the chloroform and ethanol curve.
   c) Between the ethanol and water curve.
   d) Between the water and acetic acid curve.
   e) To the right of the acetic acid curve.
4. If the temperature was 50.0° C and the atmospheric pressure was 20 kPa, which substances, if any, would boil?

Right after you read the lesson summary, you will complete Assignment 1.5. It will help you understand how to interpret vapour pressure curves. This assignment (along with all other assignments) is worth marks. You will submit it, along with all of the other assignments in this module, to the Distance Learning Unit when you have finished the module.

Lesson Summary

In this lesson, you learned how to interpret vapour pressure curves. You also used your interpolation and extrapolation skills to help you obtain data from the graph.

Assignment 1.5

Analyzing a Vapour Pressure Graph *(20 marks)*

Using the vapour pressure graph (Graph 3) from the lesson, answer the following questions. Each question is worth *1 mark*, unless otherwise indicated.

a) What is the vapour pressure of ethanol at 60° C?

____

____

b) What is the vapour pressure of acetic acid at 80° C?

____

____

c) What is the approximate vapour pressure of chloroform at 0° C?

____

____

d) What is the temperature at which the vapour pressure of ethanol is 50 kPa?

____

____

e) Which of the four substances has the highest vapour pressure?

____

____

f) Which of the four substances would evaporate fastest at room temperature? How do you know? *(2 marks)*

____

____

continued

Assignment 1.5: Analyzing a Vapour Pressure Graph (continued)

g) Which of the four substances would evaporate slowest at room temperature? How do you know? *(2 marks)*

____

____

h) Which substance has the weakest intermolecular forces? How do you know? *(2 marks)*

____

____

i) From the graph, what are the normal boiling points of the four substances? *(4 marks)*

____

____

j) The definition of vapour pressure includes the words "in a closed container." Why is it necessary to have a closed container? *(2 marks)*

____

____

k) What would the atmospheric pressure have to be in order to have ethanol boil at 20.0°C ?

____

____

l) What would the atmospheric pressure be to have acetic acid boil at 80.0° C?

____

____

m) If the temperature was 60.0°C and the atmospheric pressure was 100 kPa, which substances, if any, would boil?

____

MODULE 1 SUMMARY

Congratulations! You have reached the end of Module 1.

Submitting Your Assignments

It is now time for you to submit Assignments 1.2 to 1.5 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 1 assignments and organize your material in the following order:

- Module 1 Cover Sheet 2 (found at the end of the course Introduction)
- Assignment 1.2: Properties of Gases, Liquids, and Solids
- Assignment 1.3: Dynamic Equilibrium and Phase Changes
- Assignment 1.4: Vapour Pressure Problems
- Assignment 1.5: Analyzing a Vapour Pressure Graph

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

NOTES

Grade 11 Chemistry (30S)

Module 1: Physical Properties of Matter

Learning Activity Answer Keys

Learning Activity 1.1: Characteristics of Matter

1. For each statement, determine if it is true or false. Try to correct all statements you think are false.
   a) Liquids are easier to compress than solids.
      Answer: True
   b) Solubility is a chemical property.
      Answer: False. It is physical.
   c) Flammability is a chemical property.
      Answer: True
   d) Density is a ratio that compares the mass of an object to its volume.
      Answer: True
   e) The density of a substance changes as the substance changes state.
      Answer: True. The density can change.
   f) A chemical property does not involve a substance combining with or changing into other substances.
      Answer: False. A chemical property enables a substance to combine with or change into one or more substances.
   g) Granulated sugar is a solid form of matter with a definite shape but no definite volume.
      Answer: False. Sugar also has a definite volume.

2. Can you think of any specific examples where substances are compressed?
   Answer:
   Some examples of materials that are compressed are: tanks of oxygen gas for scuba diving and medical use, various compressed gases in asthma inhalers, compressed propane for the BBQ, compressed butane in air compressors for nail guns and pumps, and compressed liquid refrigerants in air conditioners.

3. Using your knowledge of compression, why does a tank of compressed oxygen allow scuba divers to stay underwater for long periods of time?

   Answer:

   Particles of oxygen, like most gases, are spaced far apart and are therefore easily compressed. By compressing the oxygen it is possible to put a large volume of the gas into a tank, which then allows a scuba diver to remain underwater for a long period of time.

Learning Activity 1.2: The Kinetic Molecular Theory

1. Complete each of the following statements by finding the important missing term(s).
   a) The Kinetic Molecular Theory helps us understand how the ***size***, ***motion***, and ***energy*** of particles explain the behaviour of matter.
   b) Particles that move have ***kinetic energy***.
   c) ***Gas*** particles tend to have the most kinetic energy, since they have more space between them to move.
   d) Since gas particles are separated by empty space, gases have a high degree of ***compressibility***.
   e) ***Liquids*** are also compressible, but not as easily as gases. Therefore, their degree of compressibility is less than gases.
   f) Since liquids are not easily compressed, their ***volume*** tends to be fixed.
   g) ***Solids*** also have a fixed volume, since their particles cannot be forced any closer together than they already are.
   h) Solids can be described by some of these ***properties***: high density, low kinetic energy, and high intermolecular forces.
   i) The properties of all states of matter can be explained by the ***Kinetic Molecular Theory***.

2. The following is a popular science demonstration: A few drops of water are placed in an aluminum soft drink can. The can is then placed over a flame and heated to a high temperature for a minute or so. The aluminum can is then quickly turned upside down and submerged into cold water. The can implodes and crumples before your eyes. Use the Kinetic Molecular Theory to help explain why the can collapsed. If you have access to the Internet and would like to see this demonstration, you can check out the video at <www.youtube.com/watch?v=1Efy36yxdUc>. You should also be able to find it by using the search term “can crush” in the youtube search window.

   Answer:

   When the drops of water and the air inside the soft drink can were heated, the molecules of gas gained kinetic energy and spread out, exerting pressure on the sides of the can. When the can is cooled in the cold water, the space between the gas particles decreases and less pressure is exerted on the sides of the can. Since the pressure of the air against the outside of the can is now greater than the pressure exerted by the gas molecules inside the can, the can implodes.

3. Use your own words to describe the basic assumptions of the kinetic molecular theory of gases as far as: (a) volume, (b) intermolecular forces, and (c) collisions.

   Answer:

   a) The volume of a gas particle is much less than the total volume of the gas. In other words, most of the volume of a gas is empty space.
   b) The particles of gases are so spread out that the intermolecular forces between them are negligible. In other words, gas particles do not really attract or repel each other.
   c) All collisions between gas particles are perfectly elastic. This means that there is no kinetic energy lost or gained during these collisions.

Learning Activity 1.3: Phase Changes

1. What phases are in equilibrium at a substance's melting point?

   Answer:

   The solid and liquid states.

2. Compare the evaporation of a liquid in a closed container with that of a liquid in an open container. Use the term "dynamic equilibrium" in your answer.

   Answer:

   Particles that have enough kinetic energy to overcome the intermolecular forces of attraction will move from the liquid to the vapour phase, regardless of whether they are in an open or a closed container. However, in the closed container, dynamic equilibrium will be achieved between the liquid and the vapour phases due to equal rates of evaporation and condensation.

3. Perspiration during exercise or on a hot day is an example of

   a) Condensation

   b) Sublimation

   c) Evaporation

   d) Deposition

4. Which phase changes are endothermic? Which are exothermic?

   Answer:

   Endothermic: melting, evaporation, and sublimation

   Exothermic: solidification, deposition, and condensation

5. The vaporization of a solid is also known as

   a) Condensation

   b) Deposition

   c) Evaporation

   d) Sublimation

6. Water vapour in the atmosphere will _______ to form rain.
   a) Sublime
   b) Condense
   c) Evaporate
   d) Solidify

Learning Activity 1.4: Vapour Pressure

1. Which of the following statements does not describe what occurs as a liquid boils?
   a) The temperature of the liquid increases.
   b) Energy is absorbed by the particles.
   c) The vapour pressure of the liquid is equal to atmospheric pressure.
   d) Liquid particles are entering the gaseous phase.

2. Complete the following statement: Water boils below 100º C on the top of a mountain because ***as altitude increases, atmospheric pressure decreases.***

3. In order for a liquid to boil, particles throughout the liquid must have enough ***kinetic energy*** to vaporize.

4. What would you expect the effect that increasing temperature has on vapour pressure?

   Answer:

   As temperature increases, so would the kinetic energy of the liquid particles. This would allow more vapour to form above the liquid and create a higher vapour pressure, since more particles can overcome the intermolecular forces and escape in to the gaseous phase.

5. Why is the equilibrium that exists between a liquid and its vapour in a closed container called dynamic equilibrium?

   Answer:

   Because the number of molecules evaporating are equal in number to the molecules that are condensing.

6. There is a liquid-vapour equilibrium in a sealed container. If the volume of the liquid is increased, and the container re-sealed, how will the vapour pressure be affected?
   a) The vapour pressure will increase.
   b) The vapour pressure will decrease.
   c) The vapour pressure will initially increase, and then decrease.
   d) The vapour pressure will not change.

Learning Activity 1.5: Interpolation

Using Graph 1 above, interpolate to get the following values:

1. The mass of 150 mL of this substance

   Answer: 50 g

2. The volume of 70 g of this substance

   Answer: 175 mL

3. The volume of 30 g of this substance

   Answer: 125 mL

4. The mass of 200 mL of this substance

   Answer: 85 g

5. The volume of 20 g of this substance

   Answer: 112.5 mL

Learning Activity 1.6: Extrapolation

Using Graph 2 above, interpolate to get the following values:

1. The volume of 110 g of this substance

   Answer: About 237 mL

2. The volume of 100 g of this substance

   Answer: About 225 mL

3. The mass of 275 mL of this substance

   Answer: About 138 g

Learning Activity 1.7: Interpreting a Vapour Pressure Graph

Use the vapour pressure graph (Graph 3) within the lesson to answer the following questions.

1. What would be the boiling point of water on a day when the atmospheric pressure is 95 kPa?

   Answer: 98° C

2. Ethanol is heated in a container in which there is a partial vacuum. The air pressure in the container is 25 kPa. At what temperature will the ethanol boil?

   Answer: 45° C

3. If substance "X" had a normal boiling point of 30° C, where would you expect to find the vapour pressure curve of "X"?

   a) To the left of the chloroform curve.

   b) Between the chloroform and ethanol curve.

   c) Between the ethanol and water curve.

   d) Between the water and acetic acid curve.

   e) To the right of the acetic acid curve.

4. If the temperature was 50.0° C and the atmospheric pressure was 20 kPa, which substances, if any, would boil?

   Answer: Ethanol and chloroform

Notes

Grade 11 Chemistry (30S)

Module 2: Gases and the Atmosphere

Module 2: Gases and the Atmosphere

Introduction

Most likely you have heard of natural gas, but you may be unfamiliar with its uses and where it comes from. You will begin this module by learning more about naturally occurring gases in the atmosphere and how their abundances have changed over time. This is directly tied in to air quality, an important issue about which you may already be aware. You will spend some time investigating ways air quality can be improved.

Next, you will spend the bulk of this module learning about four gas laws: Boyle's Law, Charles' Law, Gay-Lussac's Law, and the Combined Gas Law. In doing so, you will investigate the relationship between the following, as they relate to gases:

- Volume and Pressure
- Temperature and Volume
- Temperature and Pressure
- A Combination of the Above Relationships

You will end this module by applying what you have learned, by identifying gas laws at work in everyday situations.

Assignments in Module 2

When you have completed the assignments for Module 2, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 2.1	Air Quality Improvement Research
4	Assignment 2.2	Determining Significant Digits
5	Assignment 2.3	Solving Problems with Boyle's Law
6	Assignment 2.4	Investigating the Temperature-Volume Relationship
7	Assignment 2.5	Solving Problems with Charles' Law
8	Assignment 2.6	Investigating the Temperature-Pressure Relationship
9	Assignment 2.7	Problem Solving with Gay-Lussac's Law
10	Assignment 2.8	Problem Solving with the Combined Gas Law

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

Lesson 1: Gases in the Atmosphere (3 hours)

Lesson Focus

SLO C11-2-01: Identify the abundances of the naturally occurring gases in the atmosphere and examine how these abundances have changed over geologic time.
Include: oxygenation of Earth's atmosphere, the role of biota in oxygenation, changes in carbon dioxide content over time

SLO C11-2-02: Research Canadian and global initiatives to improve air quality.

Lesson Introduction

In today's modern world, few of us rely on burning whale fat as an energy source. Instead, many of us heat our homes and sometimes cook food with natural gas. In this lesson, you will learn more about naturally occurring gases in the atmosphere and how their abundances have changed over long periods of time.

What is Natural Gas?

What we call **natural gas** is actually not one gas, but a mixture of several naturally occurring gases in the atmosphere. The recipe for natural gas that we use to heat our homes, for example, is composed of the following:

- 80% methane: Methane levels have increased dramatically over the last 200 years due to the increased numbers of rice paddies, grazing animals, and landfills. It is considered to be a **greenhouse gas**, meaning it traps heat energy within the atmosphere.
- 10% ethane
- 4% propane
- 2% butane: Butane and propane can be cooled and liquefied, which allows them to be separated from the other components of natural gas. These gases are then pressurized and sold in tanks, which you might recall from Module 1.
- 4% nitrogen, helium, and other trace gases: Natural gas is a major source of helium, a gas you probably know for its use in balloons.

Combustion

In recent years, the use of natural gas has increased for a few reasons. For one, methane gas burns with a hot, clean flame. This reaction is called **combustion**. Unlike petroleum (oil pumped from the ground or tar sands), natural gas does not need to be processed and therefore does not create toxic by-products. The process of methane combustion can be represented by:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{\,(g)} + \text{heat}$$

Oxygen is necessary for this combustion reaction to occur. If there is not sufficient oxygen present, combustion will be incomplete and the flame that results will be yellow in colour. Incomplete combustion also produces carbon monoxide gas, a toxic gas that interferes with the cell's ability to carry oxygen. If there is sufficient oxygen and combustion is complete, however, the resulting flame will be blue.

Have you ever noticed the different coloured flames of a Bunsen burner? You might have observed a blue flame near the base of the burner, meaning that complete combustion of the natural gas is occurring. An orange flame from the burner means that there isn't enough oxygen to react completely with the methane.

Where Does Natural Gas Come From?

Now that you are more familiar with natural gas, you may be wondering how this useful mixture of gases is produced. Millions of years ago, marine organisms died and sank to the ocean floor. These organisms were then buried in sediment and some were fossilized. Over long periods of geologic time, *heat*, *pressure,* and *bacteria* change the residue of living organisms into natural gas and petroleum.

The combustion of natural gas is directly related to the carbon cycle, as the reaction produces both carbon dioxide and carbon monoxide. In Grade 10 Science you studied the carbon cycle. You learned that carbon dioxide is removed from the atmosphere by photosynthesis and returned to the atmosphere by cellular respiration, the decay of biological material, volcanic activity, and the burning of fossil fuels. CO_2 is a greenhouse gas, and it is considered very likely to be a major contributor to enhanced warming of the planet. An increased amount of atmospheric CO_2 (due to the increased burning of fossil fuels and deforestation) is responsible for global warming. Moreover, the amount of CO_2 that is removed from the atmosphere by photosynthesis is being reduced by the continued destruction of forest areas. This is a topic of intense debate among scientists, environmentalists, and politicians.

Life Before Oxygen

Today, the air on Earth is made mostly of nitrogen and oxygen gases, but this was not always the case. Scientists believe that before life began on Earth, the composition of the atmosphere was dramatically different than it is today. Billions of years ago the atmosphere consisted mainly of *helium* (He), *hydrogen* (H_2), *ammonia* (NH_3), *carbon dioxide* (CO_2), *water vapour* (H_2O), *methane* (CH_4), and some *nitrogen* (N_2). Earth's early atmosphere had little free oxygen and was thought to be inhospitable to life as we know it. What changed over this vast, geologic time? Origin-of-life models have generally proposed that about 1 billion years after the first primitive organisms emerged, blue-green algae appeared on the Earth. These algae converted the existing carbon dioxide and water to free oxygen (a life-supporting gas) and glucose through the process of **photosynthesis**:

$$6CO_{2\,(g)} + 6H_2O_{\,(l)} + \text{energy} \rightarrow C_6H_{12}O_6 + 6O_{2\,(g)}$$

Another important source of oxygen was the photodecomposition of water vapour by ultraviolet light according to the equation:

$$2H_2O_{\,(g)} \rightarrow 2H_{2\,(g)} + O_{2\,(g)}$$

As the amount of free oxygen increased, an ozone layer began to form filtering out ultraviolet radiation and allowing for the development of more complex species.

Let's review some key concepts regarding photosynthesis:

Question: What types of gases may have dominated the Earth before the appearance of photosynthetic organisms?

Answer: Studies suggest the following gases dominated: He, CO_2, CH_4, H_2, NH_3, N_2, and H_2O.

Question: What gas is produced by photosynthesis?

Answer: O_2

Question: What are some examples of photosynthetic organisms?

Answer: Blue-green algae, some bacteria, and most plants.

Question: Where does the energy used in photosynthesis come from?

Answer: Sunlight.

Lesson Summary

In this lesson, you learned that oxygen, a life-supporting gas, was not always as abundant as it is today. We also studied carbon dioxide, an example of a naturally occurring gas that is more abundant in the atmosphere than it used to be. This upset of the balance of the carbon cycle is having catastrophic effects on the environment.

Assignment 2.1

Air Quality Improvement Research *(10 marks)*

First, research websites and other information sources that relate to Canadian and global initiatives to improve air quality. If you do not have access to the Internet, consult your school or community librarian for guidance as to how to complete your research. You may also contact your tutor/marker for additional help.

Here are some key terms that may help in your search:

- Global Warming
- Kyoto Accord
- Climate Instability
- Climate Change
- Canadian Environmental Protection Act
- Acid Rain Strategies
- Vehicles and Fuel
- Initiatives to protect the ozone layer
- Persistent Organic Pollutants
- Canada Air Quality Health Index
- Commuter Challenge in Manitoba

Next, complete the report describing a Canadian or global initiative to improve air quality. Based on your findings, you should either promote the initiative (i.e., "The Canada Air Quality Health Index *will* help improve air quality because…") or criticize the initiative (i.e., "The Canada Air Quality Health Index will *not* help improve air quality because…"). Make sure you provide *evidence* that supports your position.

continued

NOTES

Assignment 2.1: Air Quality Improvement Research (continued)

Report: Air Quality Improvement

1. Define the problem/issue. *(3 marks, one mark per supporting point in defining the problem)*

2. Describe the strategies involved in this initiative. *(3 marks, one mark per strategy discussed)*

3. Defend of your position. *(1 mark for clearly stating your position, and 2 marks for supporting arguments)*

continued

Assignment 2.1: Air Quality Improvement Research (continued)

4. Bibliography: Use the following examples as guidelines for the completion of your bibliographical entries using MLA bibliographic reference style. *(1 mark for completion)*

   Books:

   Keyser, Daniel J., and Richard C. Sweetland, eds. *Test Critiques.* 4 vols. Kansas City, MO: TEST Corporation of America, 1988.

   Magazine Articles:

   Lamb, Douglas H., and Glenn D. Reeder. "Reliving Golden Day." *Psychology Today* (June 1986): 22–25, 28.

   Web page/Website:

   Anderson, Larry S. *Format for Update of Information Technology Plans*. 1996. www.nctp.com/html/tech_plan_format.cfm. (Accessed 13 08 10).

Lesson 2: Gas Pressure (1 hour)

Lesson Focus

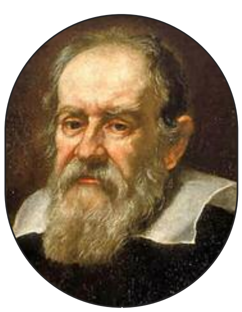
SLO C11-2-03: Examine the historical development of the measurement of pressure.
Examples: the contributions of Galileo Galilei, Evangelista Torricelle, Otto von Guericke, Blaise Pascal, Christiaan Huygens, John Dalton, Joseph Louis Gay-Lussac, Amadeo Avogadro

SLO C11-2-04: Describe the various units used to measure pressure.
Include: atmospheres (atm), kilopascals (kPa), millimetres of mercury (mmHg), millibars (mb)

Lesson Introduction

After you last had your blood pressure checked, your doctor may have given you a set of numbers such as 120/70. That blood pressure was measured in millimetres of mercury. When listening to the weather forecast, you may hear units like kilopascals being used. In this lesson, you will learn about these and other units of pressure, including how they were developed and what they actually mean.

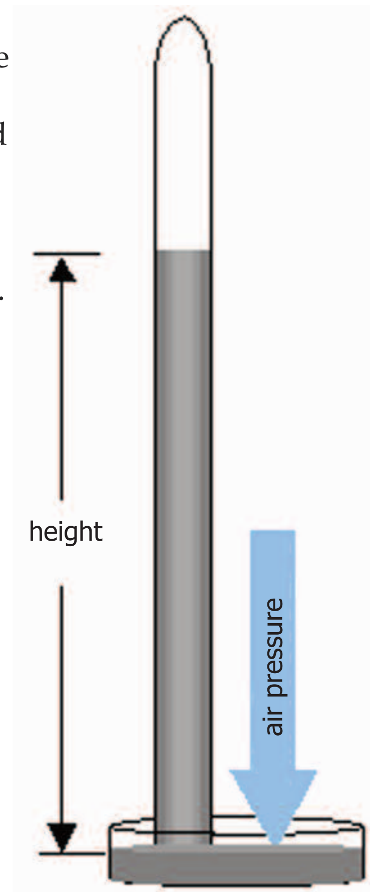
History of Measuring Pressure



Galileo Galilei (1564-1642) developed the suction pump. He used air to draw underground water up a column, similar to how a syringe draws water. He was perplexed as to why there was a limit to the height to which water could be raised. That limit was 32 feet, or about 11 metres.

In 1643, **Evangelista Torricelli** (1608-1647) developed the first **barometer**, the simple device shown below. He constructed it by inverting a closed-end tube filled with mercury into a pan of mercury at sea level. Torricelli then built upon Galileo's work by determining that the limit to the height to which Galileo's pump could draw water was due to atmospheric pressure.

The height of the column of mercury in the tube (in mmHg) represented the atmospheric pressure acting on the mercury in the pan. Torricelli determined that the height of mercury supported by atmospheric pressure at sea level was 760 mm, or 76 cm. He had performed the same experiment with water first, but found that the glass tube he had to use was too long and fragile. This application is now used in a medical device called a **sphygmomanometer**, the tool used to measure the pressure of blood on the walls of arteries when your blood pressure is taken.



Between 1643 and 1645, **Otto von Guericke** (1602-1686) made a pump that could create a vacuum so strong that a team of 16 horses could not pull two metal hemispheres apart. von Guericke reasoned that the hemispheres were held together by the mechanical force of the atmospheric pressure rather than the vacuum. The same experiment can be demonstrated by pushing two plungers together.

In 1648, **Blaise Pascal** (1623-1662) took Torricelli's barometer up and down a mountain in southern France. He discovered that the pressure of the atmosphere increased as he moved down the mountain. Sometime later, the SI unit of pressure, the 'Pascal', was named after him. He is also well known for his work in mathematics and for inventing the first calculator.

In 1661, **Christiaan Huygens** (1625-1695) developed the manometer to study the elastic forces in gases. He also developed some of the first practical vacuum pumps. Huygens, however, is better known for his work in mathematics and astronomy.

In 1801, a couple of years before publishing his atomic theory, **John Dalton** (1766-1844) stated that in a mixture of gases, the total pressure is equal to the sum of the pressure of each individual gas, as if it were in a container alone. The pressure exerted by each gas is called its partial pressure. This is known as **Dalton's Law of Partial Pressures**.

In 1808, **Joseph Louis Gay-Lussac** (1778-1850) observed the law of combining volumes. He noticed that, for example, two volumes of hydrogen combined with one volume of oxygen to form two volumes of water. He is also well known for his passion for hot air ballooning, which helped him to conclude that there is a direct relationship between the pressure and temperature of a gas.

After studying the work of Gay-Lussac and others, **Amadeo Avogadro** (1776-1856) published what is known as **Avogadro's Hypothesis** in 1811. His hypothesis stated that a sample of any gas at the same temperature and pressure will contain the same number of particles. Unfortunately, because Avogadro did not perform his own experiments, his hypothesis was ignored for about 50 years.

Units of Pressure

You may recall from the last module that gas particles are in constant random motion. As such, the particles collide with the sides of the container in which they are held, in what is called an elastic collision. The force per unit area due to these collisions is called gas pressure. There are several units used to represent this pressure.

The unit **atmosphere (atm)** was derived from standard atmospheric pressure at sea level. One atmosphere is equal to 760 mmHg, or 101.325 kPa. Two atmospheres are twice standard atmospheric pressure, and so on.

The SI unit for pressure is the **Pascal (Pa)**, named after Blaise Pascal. The Pascal is defined as one Newton of force per square metre. One Pascal is equal to the pressure exerted by a stamp on an envelope. As this is a rather small unit, when measuring gas pressure we use **kilopascals** (equal to 1000 Pascals) instead. You have probably heard of the kilopascal while listening to weather reports. Standard atmospheric pressure is 101.325 kPa.

Millimetres of mercury (mmHg) is not a common unit used outside the laboratory today; however, many barometers found in homes use both millimetres of mercury as well as another unit like kilopascals. In the United States, air pressure is often reported in terms of inches of mercury.

The **millibar** is a meteorological unit of atmospheric pressure. You may remember seeing this unit used on some weather maps in Grade 10 Science. Again, this is a rather small unit, so scientists tend to use bars. 1013 mbars is the same as 1.013 bar. They are both equal to standard atmospheric pressure, or 1 atmosphere.

One other common pressure unit is **pounds per square inch (psi)**. This is an Imperial unit of pressure. This unit is often encountered when filling car and bicycle tires. One kilopascal is equal to 0.145 psi.

Here is a summary of the units that are used to represent pressure:

$$1 \text{ atm} = 101\,325 \text{ Pa} = 101.325 \text{ kPa} = 760 \text{ mmHg} = 14.69 \text{ psi}$$

To help you review what you have just learned, complete Learning Activity 2.1. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 2.1

Measuring Pressure

Part A: Complete the following multiple choice questions by circling the best possible response.

1. This scientist is credited with inventing the barometer.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
2. This scientist developed the manometer and the first useful vacuum pump.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
3. This scientist discovered that water could only be raised to a certain height in a column.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
4. This scientist discovered that the pressure of the atmosphere changes according to elevation.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens

continued

5. This scientist theorized that a sample of any gas at the same temperature and pressure will contain the same number of particles.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton
6. This scientist created a very strong vacuum that furthered our understanding of atmospheric pressure.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton
7. This scientist experimented with combining volumes of gases and created a gas law based on his findings.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton
8. This scientist stated that the total pressure of gases in a container is equal to the sum of the pressures of each individual gas.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton

continued

Learning Activity 2.1 (continued)

Part B: Complete the following statements with the correct missing term(s).

1. A measurement unit of pressure is named after ____________ .
2. A device used to measure air pressure is called a ____________ .
3. Atmospheric pressure ____________ as elevation increases.
4. In fair weather and at sea level, the atmospheric pressure is sufficient to support a column of mercury about ________________________ high.
5. How many atmospheres of pressure are required to support 760 mm of mercury in a barometer? ____________
6. The SI unit for pressure is the ____________ , which represents a small amount of pressure.
7. Normal atmospheric pressure is ____________ kPa.
8. The __________________________ is the unit of pressure commonly used in meteorology.
9. ____________ is defined as force per unit area.
10. Gas pressure is usually measured with an instrument called a ____________ .

Lesson Summary

Gas pressure is the result of billions of rapidly moving gas particles colliding with an object. In this lesson, you reviewed the people who improved our understanding of gas pressure and the units that are used to measure pressure. In the next lesson, you will investigate how a change in pressure affects the volume of a contained gas.

NOTES

Lesson 3: Gas Pressure and Volume, Part A (1.5 hours)

Lesson Focus

SLO C11-2-05A: Experiment to develop the relationship between the pressure and volume of a gas using visual and graphical representations.
Include: historical contributions of Robert Boyle

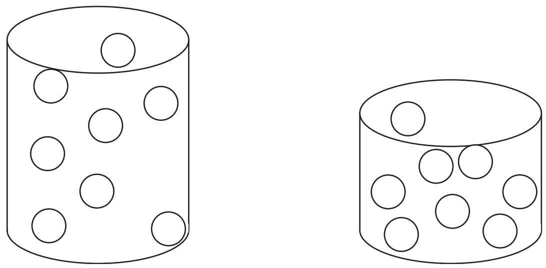
Lesson Introduction

As you breathe, there is a volume-pressure relationship at work in your lungs. When you breathe in, your diaphragm drops, which allows the volume of your lungs to increase. At the same time, the air pressure in your lungs decreases. Since the air pressure outside of your lungs is now greater than the pressure inside your lungs, air flows into your lungs easily.

As you breathe out, your diaphragm rises, decreasing the volume of your lungs so that the pressure inside increases. This forces air to move outside of your body. This is one example of Boyle's Law at work. In this lesson, you will investigate the relationship between gas volume and pressure in more detail.

The Work of Robert Boyle

You have likely noticed **Boyle's Law** when you squeezed a balloon or sat on an air mattress. Decreasing the volume of a sample of gas, while keeping the temperature constant, results in the same number of molecules being squeezed into a smaller space. This will increase the number collisions of the gas molecules with the sides of the container, and thus the gas pressure.



Imagine trying to fit a room full of people into a closet. As people tried to move around, they would continually bump into the other people in the small area. This is just like decreasing the volume of a sample of gas at a constant temperature—an increased number of collisions between gas particles in a smaller area results in increased pressure.

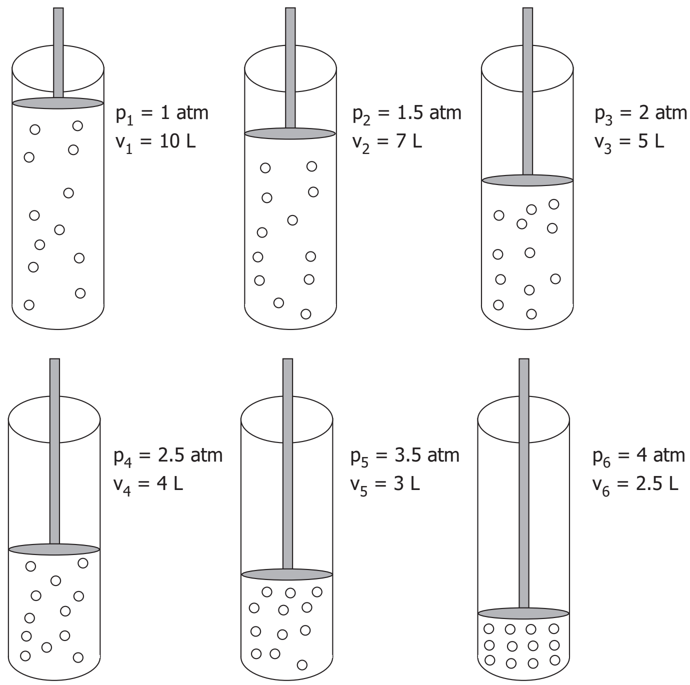
Now, take the people out of the closet and put them back in the room. It would be less likely that people would bump into each other as they moved around, because there would be more space between them. This is just like increasing the volume of a sample of gas at a constant temperature. Increasing the volume reduces the number of collisions between the gas molecules and the sides of their container.

To help you review what you have just learned, complete Learning Activity 2.2. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 2.2

Investigating the Pressure-Volume Relationship

Below you will see six cylinders, each one containing the same type of gas at the same temperature. Each cylinder has a plunger that is compressing the gas to a certain volume, thus creating a certain amount of pressure in each cylinder. After you observe these six cylinders, use the data table to record the volume and pressure of each one. The first set of data has been recorded for you.



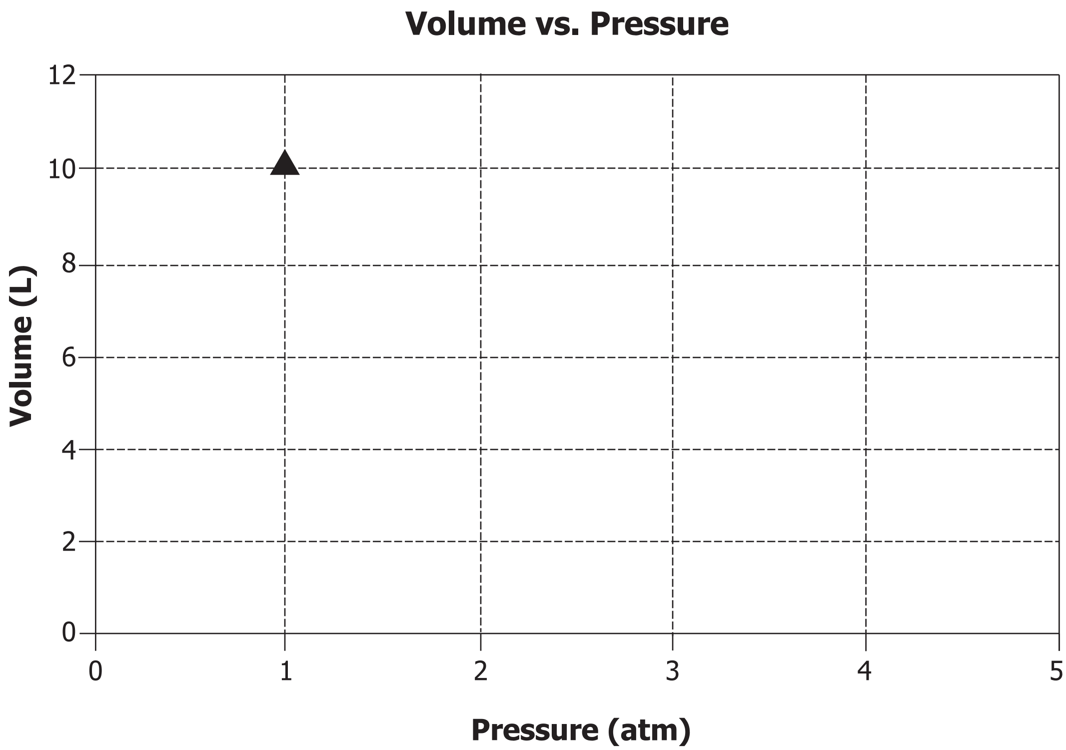
Pressure (atm)	Volume (L)
1	10

continued

Learning Activity 2.2 (continued)

Using the data from your table, you can now complete the Pressure-Volume graph below. We will refer to this as Graph 1. The first point has been plotted for you. Once all six points are plotted, draw a smooth curve linking all of the points together.

Graph 1



continued

Learning Activity 2.2 (continued)

Use the Pressure-Volume graph (Graph 1) to answer the following questions:

1. What is the relationship between the pressure and the volume of a gas? Give an example from the graph.
2. From the graph, what would the volume of gas be if the pressure were 3.5 atm?
3. From the graph, what would the volume of gas be if the pressure were 1.5 atm?
4. From the graph, what would the gas pressure be if the volume were 6.0 L?
5. From the graph, what would the gas pressure be if the volume were 3.0 L?
6. What would the volume be if the pressure were increased to 5 atm? Hint: You must extrapolate the curve before you can answer the question.
7. Why is it dangerous to puncture an aerosol can?

What Is an Inverse Relationship?

You may remember learning about inverse relationships in your math class. Since it is important to understand this type of relationship before we investigate Boyle's Law, let's do some review. If one variable increases while the other variable decreases, then there is an inverse relationship between those variables. An example of this would be the number of people helping versus how long it takes to paint a gym mural. As more people help, the time it takes decreases. The opposite of an inverse relationship is a direct relationship. In this case, as one variable increases so does the other one. An example of this would be the weight of a diamond versus how much it costs. Generally speaking, as a diamond increases in weight, its cost also increases.

In an inverse relationship, the product of the two variable quantities is constant. This means the product of pressure and volume at any two sets of pressure-volume conditions is always constant as long as the temperature is held constant. For example, if you look at Graph 1 above, we can illustrate this principle using the first and third points of data that were plotted.

$$(P)\ (V) = (10\ L)\ (1\ atm) = 10\ L{*}atm = k$$

$$(P)\ (V) = (5\ L)\ (2\ atm) = 10\ L{*}atm = k$$

To help you review what you have just learned, complete Learning Activity 2.3. Like all other learning activities in the course, it will help you prepare for your assignments and examinations. Check your answers against those provided in the Learning Activity Answer Keys found at the end of this module.

Learning Activity 2.3

The Relationship between Volume and Pressure

1. Use the graph in the lesson to describe the inverse relationship between the volume at 1 atm and 2 atm.
2. You release a helium-filled balloon into the sky. If the temperature remains constant, what will happen to the volume and pressure of the helium as the balloon rises higher in the sky?
3. If a gas is compressed from 2 L to 1 L and the temperature remains constant, what will happen to the pressure? Why?
4. Use your own words to explain Boyle's Law.

Lesson Summary

Knowing more about Robert Boyle and his gas law, you may recognize the relationship between the volume and pressure of gas at work in your everyday life. Now you can explain why squeezing a foam ball is easy at the beginning (when volume is high and pressure is small), but gets increasingly difficult as the volume of the ball is reduced (more pressure is required to compress the ball any further). The law is also the basis for the popular Cartesian diver and marshmallow in a vacuum chemistry demonstrations. If you have Internet access and want to learn more about these applications of Boyle's Law, check out the following websites:

- www.metacafe.com/watch/877495/cartesian_diver/ and
- www.metacafe.com/watch/898336/in_space_without_a_ spacesuit/

In the future lessons, you will continue your investigation of this relationship.

NOTES

Lesson 4: Working with Significant Figures (1 hour)

Lesson Focus

After completing this lesson, you will be able to

- ❑ understand the importance of expressing calculation answers to the correct number of significant figures
- ❑ perform calculations that result in the correct number of significant figures

Lesson Introduction

If we do calculations on our calculator using numbers representing quantities that we have measured, how do we know where to round off our answer? Should we keep all of the digits in the calculator display? If not, where should we round off the answer?

The Purpose of Significant Digits

When we count, we use exact numbers. If we count people, we might have exactly four (4) people. There is no uncertainty about this number of people.

Measurements using an interval scale, like that found on a ruler, do have some uncertainty or estimation involved when we make a measurement. For example, you might weigh a bag of oranges at the grocery store on a scale calibrated in 0.1 kg intervals and you notice that your oranges weigh between 1.8 and 1.9 kg. You then estimate their weight to be around 1.85 kg. The number in your estimated measurement has three digits, the first two being known with certainty, but the third digit being somewhat uncertain. Perhaps your oranges actually weigh 1.87 kg, or 1.88 kg. The significant figures in this, and any other measurement, include all digits that are known, plus the last digit that is *estimated*.

Significant digits (a.k.a. **significant figures**) tell us something about how the measurement was made. A better instrument allows us to make better measurements, so that we can record a greater number of significant digits in reporting a value. A metre stick would only allow us to record the thickness of a hair as 0.1 mm, while a micrometer (a much better measuring instrument) allows us to record the thickness as 0.137 mm. This second measurement is closer to the true value and is therefore more accurate.

Recording a Measurement Using Significant Digits

When recording a measurement, we must include every digit that is absolutely certain plus the first digit that must be estimated. This is the definition of a significant digit.

Significant digits are part of a measurement. For example, suppose you measure the length of a table with a 1 m ruler calibrated to 10 centimetres. You notice the table has a length of 2 metres, plus 60 centimetres, plus a little bit more. The table is longer than 2.6 metres, but less than 2.7 metres. You might then record the length of the table as 2.64 metres. The first two digits are known exactly and the third digit is estimated, or a guess. The measurement has three significant digits.

A measurement of the same table with a ruler calibrated to centimetres could yield a value of 2.642 metres. The better instrument provides a greater number of significant digits when measuring the same object.

The final significant digit will always be one unit smaller than the calibration of the measuring instrument. For example, the first measurement above was recorded to the nearest centimetre from a ruler that was calibrated to the nearest 10 centimetres. The second measurement was recorded to the nearest tenth of a centimetre from a ruler calibrated to the nearest centimetre.

Zeroes (0) are to be recorded only if they are significant. For instance, if a table was measured and the end of the table coincided exactly with a mark on the ruler (for example, 2.6 metres), a zero must be recorded. If the ruler is calibrated to 10 centimetres, a "0" should be recorded as the next smaller unit. In this case, the measurement would be 2.60 metres. If the metre stick were calibrated in centimetre units, then zeroes would be recorded until the one millimetre place value. Here, the measurement would be 2.600 metres.

Rules for Significant Digits

The following rules are used to determine the number of significant digits (a.k.a. figures) in a given measurement. You may have already learned these rules in another science or math course.

1. *All non-zero digits are significant.*

   Examples:

   374 (three significant digits)

   8.1 (two significant digits)

2. *All zeroes between non-zero digits are significant.*

   Examples:

   50407 (five significant digits)

   8.001 (four significant digits)

3. *Leading zeroes in a decimal value are not significant.*

   Examples:

   0.54 (two significant digits)

   0.0098 (two significant digits)

4. *Trailing zeroes are significant if they are to the right of a decimal point.*

   For example, 14.0 has three significant figures. Note that the zero in the tenth's place is considered significant. All trailing zeros in the decimal portion are considered significant. A common error is to think that 14 and 14.0 are the same thing. THEY ARE NOT. 14.0 is ten times more precise than 14. The two numbers have the same value, but they convey different meanings about how precise they are.

   Examples:

   2370 (three significant digits)

   16000 (two significant digits)

   160.0 (four significant digits)

5. *In numbers greater than 1, trailing zeroes are not significant unless stated so.*

   Examples:

   37000 (5 significant digits, because the decimal point identifies that the trailing zeroes are significant)

   37000 (2 significant digits, because the lack of a decimal point identifies that trailing zeroes are not significant)

To show when trailing zeroes are significant, we use **scientific notation**. All zeroes present when a number is written in scientific notation are significant.

Examples:

37000 with 3 significant digits would be 3.70×10^{-4}

37000 with 4 significant digits would be 3.700×10^{-4}

37000 with 5 significant digits would be 3.7000×10^{-4}

Addition and Subtraction Using Significant Figures

In mathematical operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the least precise operation. In other words, a chain is no stronger than its weakest link. If you are running on a relay team, the slowest member of the team will dictate how fast your team can finish. Likewise, an answer is no more precise than the least precise number used to get the answer. Your answer cannot be MORE precise than the least precise measurement.

When adding or subtracting numbers, your answer should be rounded to the same number of decimal places (not digits) as the measurement with the LEAST number of decimal places. Here is an example:

$3.461728 + 14.91 + 0.980001 + 5.2631$

Step 1: Count the number of significant figures in the decimal portion of each number in the problem. (The digits to the left of the decimal place are not used to determine the number of decimal places in the final answer.) Since 14.91 has the least number of digits to the right of the decimal point, the answer must also be rounded to two decimal places.

Step 2: Add or subtract in the normal fashion. Your answer would be 24.614829 (before rounding).

Step 3: Round the answer to the LEAST number of places as determined by the numbers in the problem. We determined that our answer should reflect two decimal places, and so 24.61 is the correct amount of significant figures.

An Alternative Method for Adding and Subtracting

You can also add a set of numbers in column form. If you were asked to add 23.1, 4.77, 125.39, and 3.581, you would start by adding as usual to find the preliminary answer (156.841).

Next, you would draw a vertical line after the last significant digit in the number with the least precision (largest guess). Then, you would round off the answer to the last place value BEFORE the vertical line.

23.1		guess in	0.1	← least precision (largest guess)
4.7	7	guess in	0.01	
125.3	9	guess in	0.01	
3.5	81	guess in	0.001	
156.841				

Round off in the 0.1 place value, the place value of the least precision (largest guess) from the numbers in the question. The answer to the question is then 156.8.

Since you already know where the decimal is, drop the trailing digits and do not replace them with zeroes.

Multiplication and Division Using Significant Figures

When multiplying and dividing, the LEAST total number of significant figures in any number of the problem determines the number of significant figures in the final answer. This means you MUST know how to recognize significant figures in order to use this rule.

Example 1

2.5 x 3.42

The answer to this problem would be 8.6, rounded from the calculator reading of 8.55. Why? 2.5 has two significant figures while 3.42 has three. Two significant figures is less precise than three, so the answer can only have two significant figures.

Example 2

How many significant figures will there be in the answer to

3.10 x 4.520?

You may have said two. This is too few. A common error is to look at a number like 3.10 and think it has two significant figures. The zero in the hundredth's place is recognized as significant. Therefore, 3.10 has three significant figures, while 4.520 has *four* significant figures. The answer should only have three significant figures.

Example 3

$(4.52 \times 10^{-4}) \div (3.980 \times 10^{-6})$.

How many significant figures should there be in the answer?

Answer: Three.

Which number decided this?

Answer: 4.52×10^{-4}, since it was the less precise of the two numbers.

The Atlantic-Pacific Rule

Sometimes there are helpful little tools that can help you remember key information. When trying to keep significant figure rules straight, the Atlantic-Pacific rule may help (assuming, of course, that you know where these two oceans are located). The rule works this way:

If a decimal point is *present*, ignore zeros on the *Pacific* (left) side of a number. If the decimal point is *absent*, ignore zeros on the *Atlantic* (right) side of a number. Everything else is significant.

For students who do not live in North America, or if you are unsure of your geography, you may prefer the following way to say the same thing:

- Ignore leading zeros.
- Ignore trailing zeros, unless they come after a decimal point.
- Everything else is significant.

Here are the general *rules for rounding*:

If the number you are rounding is followed by 5, 6, 7, 8, or 9, round the number *up*. For example, 389 rounded to the nearest hundred is 400.

If the number you are rounding is followed by 0, 1, 2, 3, or 4, round the number *down*. For example, 23 rounded to the nearest ten is 20.

Note that statisticians prefer to round 5 to the nearest even number. As a result, about half of the time 5 will be rounded up, and about half of the time it will be rounded down. For example, 24.5 rounded to the nearest whole number would be 24—it would be rounded down. And 75.5 rounded to the nearest whole number would be 76—it would be rounded up.

Learning Activity 2.4

Working with Significant Figures

1. Determine the number of significant digits in each of the following numbers.

5.897		8.000		10001	
0.333		8.001		0.008000	
7		0.009		947.000	
10000		12000		10000.0	
10321		55040		375000	

2. Complete the following questions respecting the correct number of significant figures in your answers.
   a) 23.1 + 4.77 + 125.39 + 3.581
   b) 22.101 – 0.9307
   c) 2.33 x 6.085 x 2.1
   d) 8432 ÷ 12.5

Lesson Summary

Significant digits are used when we make measurements or perform calculations with measured quantities. Counting does not involve any guesswork or estimating, so such numbers are said to be exact. When we measure using an interval scale, we record the digits we know for sure plus one estimated digit. This is our definition of significant digits.

Notes

Assignment 2.2

Determining Significant Digits *(10 marks)*

1. State the number of significant digits in each measurement. *(0.5 marks each = 6 marks)*

2509 m		7.62 km	
0.00055 m		0.0670 m	
5.060×10^{5} m		9.0000×10^{-5} m	
240 m		2.4 m	
2400 m		2400.0 m	
0.005050 m		50 m	

2. Complete the following problems, respecting the correct number of significant figures in your answers. *(4 marks)*

   a) $0.04216 - 0.0004134$

   ____

   b) $564.321 - 264.321$

   ____

   c) 35.2×64.89

   ____

   d) $2.4526 \div 8.4$

   ____

NOTES

Lesson 5: Gas Pressure and Volume, Part B (2 hours)

Lesson Focus

SLO C11-2-05B: Experiment to develop the relationship between the pressure and volume of a gas using numeric representation.
Include: historical contributions of Robert Boyle

Lesson Introduction

Boyle's Law is an example of a relationship where one variable increases as the other variable decreases. In Lesson 3 you examined this type of a relationship between the pressure and volume of a gas. You observed that as the pressure of a gas increased, the volume of the gas decreased, and vice versa. In this lesson you will use mathematics to help you understand this inverse relationship.

Boyle's Law

Robert Boyle (1627-1691) experimented to demonstrate an inverse relationship between the volume and pressure of a gas at a constant temperature. In doing so, he found that doubling the pressure of a gas decreased its volume by about one-half. This is the inverse relationship you investigated in Lesson 3. When you examined Graph 1 from Lesson 3, you observed that the product of pressure and volume at any two sets of pressure-volume conditions is always constant at a constant temperature. You proved this by comparing the first and third points of data plotted on Graph 1. In both instances, the product of pressure and volume was 10 L∗atm:

$$(P)(V) = (10\ L)(1\ atm) = 10\ L*atm = k$$

$$(P)(V) = (5\ L)(2\ atm) = 10\ L*atm = k$$

Robert Boyle also expressed this inverse relationship as an equation. This *mathematical relationship*, known as the **Boyle's Law equation**, can be expressed as follows:

$$P_1V_1 = P_2V_2$$

In the above expression, P_1 and V_1 represent the initial pressure and initial volume of a sample of gas, while P_2 and V_2 represent the pressure and volume of the same sample of gas under new conditions. Boyle's Law assumes that the temperature does not change, but rather stays constant throughout.

If you know any three of the four variables from the Boyle's Law equation, you can solve for the missing variable. Use the steps below to help you;

1. *Analyze the data,* listing both the known and unknown variables.
2. *Set up the two possible ratios.*
3. *Evaluate* if your answer is logical.

Example 1

If 3 L of gas is initially at a pressure of 1 atm, what would be the new pressure to cause the volume of the gas to become 0.5 L?

Ask yourself, "What happens to the volume?" It decreases.

According to Boyle's Law, how will the pressure change with decreased volume? Decreasing volume results in an increase in pressure. Use the same sequence of steps from above to solve this problem.

Step 1: A*nalyze your data.*

$P_1 = 1$ atm $\qquad P_2 = ?$

$V_1 = 3$ L $\qquad V_2 = 0.5$ L

Step 2: *Set up the two possible ratios.* Looking at the information provided, you can see that the volume of gas has decreased. Decreasing the volume from 3 L to 0.5 L should result in an increase in pressure. There are two possible volume ratios:

$$\frac{3\text{ L}}{0.5\text{ L}} \text{ or } \frac{0.5\text{ L}}{3\text{ L}}$$

Step 3: *Evaluate.* Does your answer make sense? Only one of these ratios, when multiplied by the pressure, will result in a higher pressure—the first ratio.

$$\text{new pressure} = 1\text{ atm} \times \left(\frac{3\text{ L}}{0.5\text{ L}}\right) = 6\text{ atm}$$

Therefore, P_2 = 6 atm of pressure. This means that 6 atm of pressure will change 3 L of a gas at 1 atm to 0.5 L.

You can verify your answer by using the Boyle's Law equation:

$P_1V_1 = P_2V_2$

(3 L)(1 atm) = (x)(0.5 L)

Using your cross-multiplication skills, you can solve for the unknown variable, x.

Here is another example:

Example 2

A syringe contains 20 mL of a gas at 100 kPa. The pressure in the syringe is changed to 25 kPa. What is the new volume of the gas?

Ask yourself, "What happens to the pressure?" In this case, it decreases.

According to Boyle's Law, how will the volume change with a decrease in pressure? Decreasing pressure results in an increase in volume.

Use the same sequence of steps from the first example to solve this problem.

Step 1: *Analyze your data.*

P_1 = 100 kPa $\qquad$ P_2 = 25 kPa

V_1 = 20 mL $\qquad$ V_2 = ?

Step 2: *Set up the two possible ratios.* There are two possible pressure ratios:

$$\frac{100 \text{ kPa}}{25 \text{ kPa}} \text{ or } \frac{25 \text{ kPa}}{100 \text{ kPa}}$$

Step 3: *Evaluate your answer.* Only one of these ratios, when multiplied by the volume, will result in an increased volume—the first ratio.

$$\text{new volume} = 20 \text{ mL} \times \left(\frac{100 \text{ kPa}}{25 \text{ kPa}}\right) = 80 \text{ mL}$$

Therefore, V_2 = 80 mL at the new pressure of 25 kPa.

Now let's check using the Boyle's Law equation:

$P_1V_1 = P_2V_2$

(100 kPa)(20 mL) = (25 kPa)(x)

Using your cross-multiplication skills, you can solve for the unknown variable, x. You should be able to verify the new volume to be 80 mL.

Learning Activity 2.5

Using Boyle's Law

1. Use Boyle's Law to calculate the following. Show your work and always use significant figures for your answers.
   a) V_1 = 2.0 L, P_1 = 0.82 atm, V_2 = 1.0 L, P_2 = ?
   b) V_1 = 0.55 L, P_1 = 740 mm Hg, V_2 = 0.80 L, P_2 = ?
   c) V_1 = 4.0 L, P_1 = 210 kPa, V_2 = 2.5 L, P_2 = ?
2. The volume of a gas at 98 kPa is 250 mL. If the pressure is increased to 180 kPa, what will the new volume be?
3. A gas is placed into a syringe until the pressure is 45.0 kPa. What is the new pressure if
   a) the volume in the syringe is doubled?
   b) the volume in the syringe is tripled?
   c) the volume is one-third its original volume?

Lesson Summary

In this lesson, you studied the effect of changing the pressure and volume of a gas at constant temperature. In the next lesson, you will investigate the effect of heating a gas at constant pressure.

Assignment 2.3

Solving Problems with Boyle's Law *(21 marks)*

1. 100.0 mL of gas is placed into a syringe. What is the new volume if
   a) the pressure is doubled? *(1 mark)*

   b) the pressure is tripled? *(1 mark)*

   c) the pressure is one-quarter the original pressure? *(1 mark)*

continued

Assignment 2.3: Solving Problems with Boyle's Law (continued)

2. Change the following from the initial conditions to the new conditions using the ratio method. First, make a verbal analysis of the data based on your knowledge of Boyle's Law. Next, solve the problem showing all your work. Be sure to include the appropriate units for your solutions. *(3 marks each; 1 mark for prediction, 1 mark for work, 1 mark for correct answer = 18 marks)*

   a) 100.0 mL of oxygen gas at 10.50 kPa is changed to 9.91 kPa. What is the new volume?

   b) 50.0 cm3 of helium at 97.3 kPa is changed to 102.5 kPa. What is the new volume?

continued

Assignment 2.3: Solving Problems with Boyle's Law (continued)

c) 25.0 mL of nitrogen at 0.990 atm is changed to 0.751 atm. What is the new volume?

d) 0.550 L of hydrogen at 745 torr is changed to 0.700 L. What is the new pressure?

continued

Assignment 2.3: Solving Problems with Boyle's Law (continued)

e) 1.32 L of oxygen at 1.40 atm is changed to 0.705 L. What is the new pressure?

f) 525 mL of neon at 49.3 kPa is changed to 845 mL. What is the new pressure?

Lesson 6: Temperature and Volume, Part A (1.5 hours)

Lesson Focus

SLO C11-2-06A: Experiment to develop the relationship between the volume and temperature of a gas using visual and graphical representations.
Include: historical contributions of Jacques Charles, the determination of absolute zero, the Kelvin temperature scale

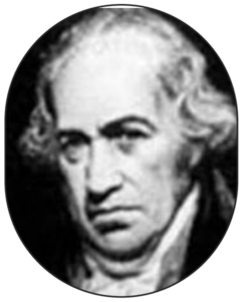
Lesson Introduction

Jacques Charles discovered the relationship between gas temperature and volume while ballooning, a finding that sparked King Louis XVI to grant Charles his own laboratory. You may have observed the same temperature-volume relationship while baking bread or a cake. As the dough heats in the oven, the carbon dioxide gas produced by the yeast and baking powder expands. The expanding volume of gas causes the cake or bread to expand (or rise) too, much like a hot air balloon. In this lesson, you will investigate this direct relationship between gas temperature and volume.

How Gas Temperature and Volume Are Related

Since the dawn of civilization, people have experienced the relationship between temperature and volume. After all, heating a gas in a cake or bread makes the cake or bread rise more!

In addition to the relationship between pressure and volume, Robert Boyle had also studied the relationship between temperature and volume, but he failed to discover what we call Charles' Law because a temperature scale did not exist. The precise mathematical relationship between temperature and volume was not described until 1699 when it was discovered by the French physicist **Guillaume Amontons** (1663-1705) after he developed a thermometer based on the increasing volume of a gas (instead of a liquid) with an increase in temperature.



In 1714, **Daniel Fahrenheit** (1686-1736) invented the mercury thermometer and the temperature scale that is named after him. He proposed 0 to be the coldest temperature in Western Europe and 100 to be the highest temperature. This made the freezing point of water 32°F and the boiling point of water 212°F.

The work of Amontons and Fahrenheit paved the way for **Jacques Charles** (1746-1823), a French scientist, who only had a basic knowledge of mathematics and almost no science education. He had only become interested in non-mathematical, experimental physics when Benjamin Franklin visited Paris in 1779 as an ambassador for the new United States. Inspired by the popularity of hot-air balloons in his time, Charles investigated the expansion rates of different gases due to temperature changes.

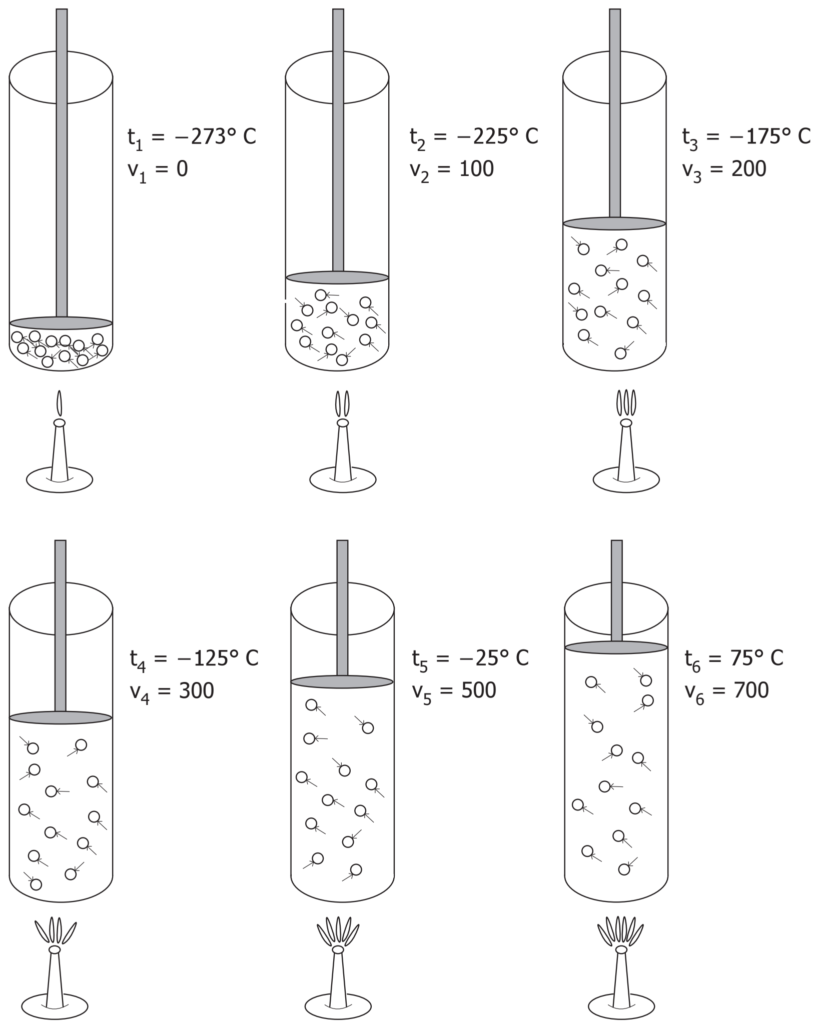
In his experiments, Charles observed that as temperature increases, so does the volume of a sample of gas when the pressure is held constant. This relationship is known as **Charles' Law**. The kinetic-molecular theory can shed some light on this relationship. You already know that as a gas is heated, the speed, force, and frequency of particle collisions increase. This would usually increase the gas pressure as well, but by increasing the volume of gas, particles have to travel farther before colliding with the container. This increase in volume allows gas pressure to remain constant.

Today, chemistry teachers and science magicians take advantage of Charles' Law when they use a fire syringe (Check out www.teachertube.com and type in the search terms *fire syringe termodynamics* to see how it works) or demonstrate what happens to the air in a latex glove when it is exposed to liquid nitrogen.

Assignment 2.4

Investigating the Temperature-Volume Relationship
(20 marks)

Below you will see six cylinders, each one containing the same type of gas at the same pressure. Each cylinder of gas is heated to a different temperature, as indicated. When the cylinder is heated, the kinetic energy of the gas particles also increases. The volume of the gas increases, forcing the plunger upwards. The distance the plunger is forced upward is a measure of the increase in gas volume due to heating the gas.



continued

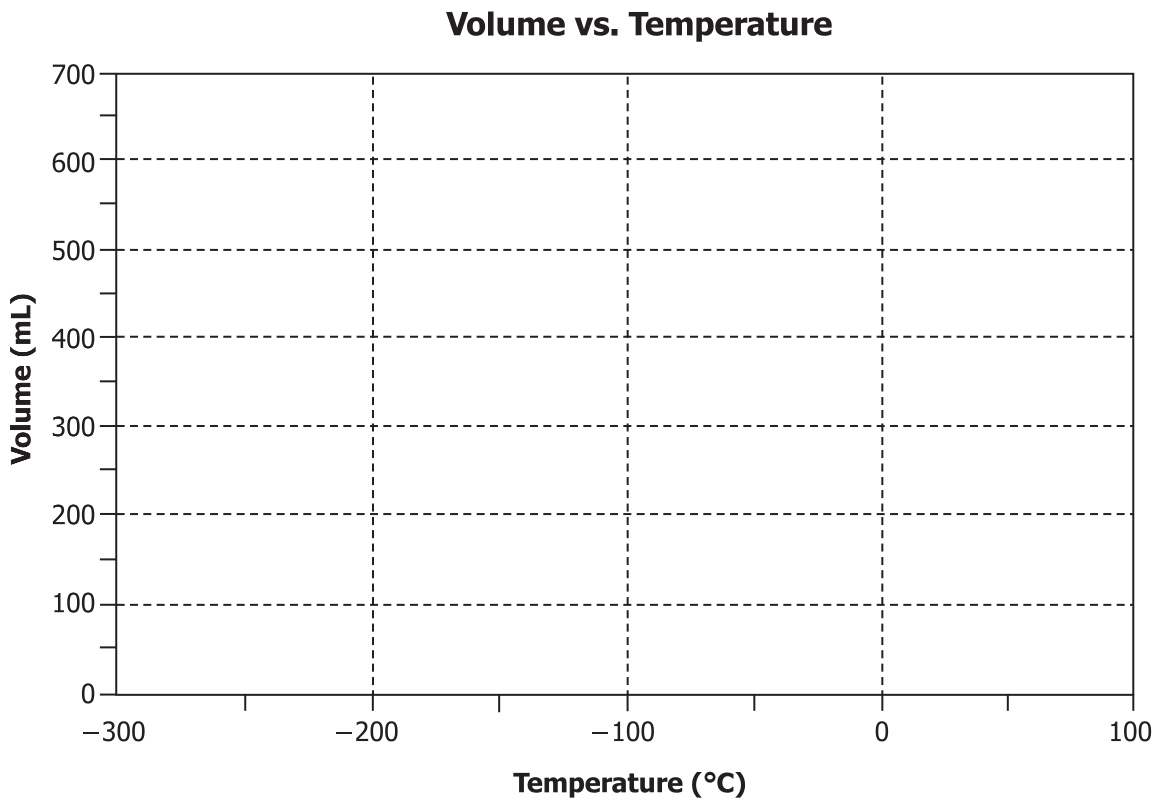
Assignment 2.4: Investigating the Temperature-Volume Relationship (continued)

After studying these six cylinders, use the data table to record the volume and temperature of each one. *(1 mark per set of data = 6 marks)*

Temperature (°C)	Volume (mL)

Using the data table, plot the six sets of points on the blank graph provided. *(1 mark per point = 6 marks)*

We will refer to this as Graph 2. Don't forget to draw the line of best fit. *(1 mark)*



continued

Assignment 2.4: Investigating the Temperature-Volume Relationship (continued)

Use Graph 2 to answer the following questions.

1. What do you notice about the temperature at zero volume on the graph? *(1 mark)*

2. What do you notice about the volume at zero temperature on the graph? *(1 mark)*

3. Is the relationship between the volume and temperature of a gas inverse? How do you know? *(2 marks)*

4. What happens to the volume of a gas as its temperature increases? *(1 mark)*

5. Use your knowledge of the kinetic molecular theory to explain the relationship between gas volume and temperature. *(2 marks)*

NOTES

Lesson 7: Temperature and Volume, Part B (1.5 hours)

Lesson Focus

SLO C11-2-06B: Experiment to develop the relationship between the volume and temperature of a gas using numeric representation.
Include: historical contributions of Jacques Charles, the determination of absolute zero, the Kelvin temperature scale

Lesson Introduction

In the previous lesson, you examined the direct relationship between the temperature and volume of a gas. You observed that as temperature increased, the volume of a contained gas also increased. In this lesson you will use mathematics to help further your understanding of Charles' Law.

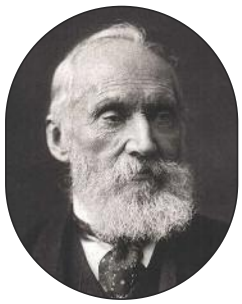
Measuring Temperature

Twenty-eight years after Daniel Fahrenheit invented the mercury thermometer and the temperature scale that is named after him, **Anders Celsius** (1701-1744) invented the centigrade (Celsius) temperature scale that is often named after him. He used the freezing point of water as 0°C and the boiling point of water as 100°C.

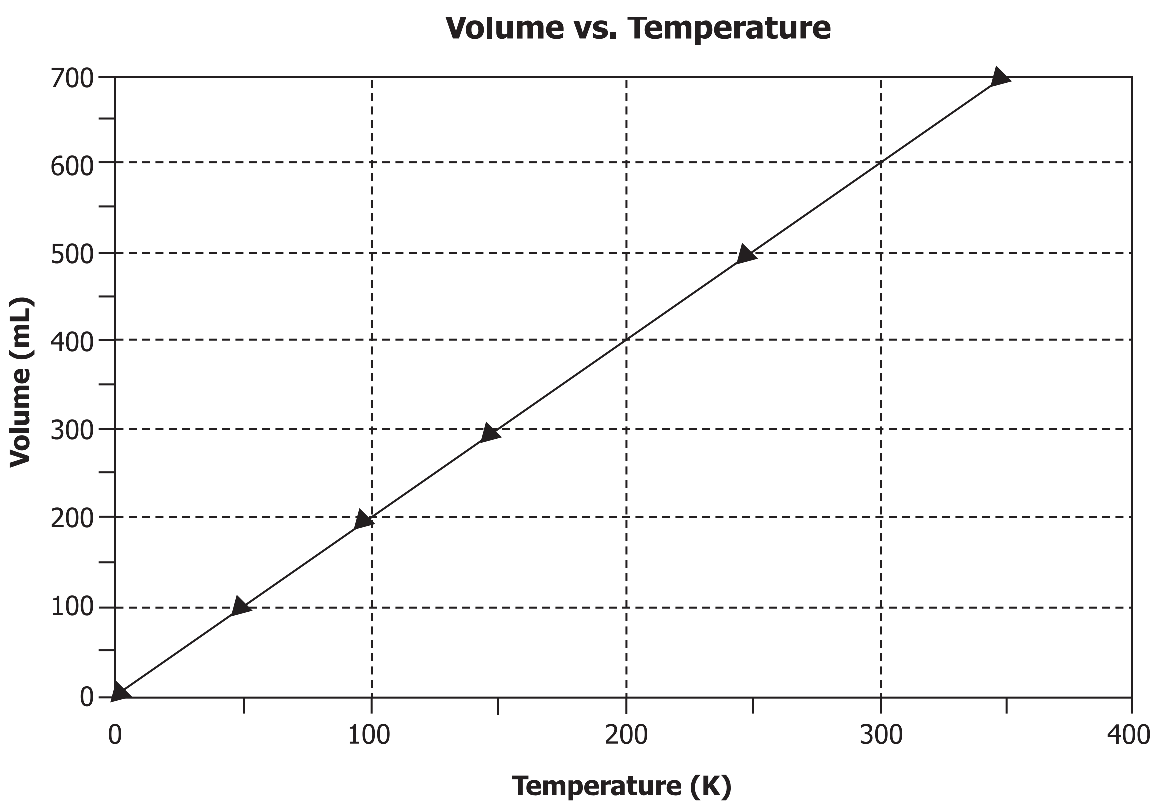
When Jacques Charles studied the effect of temperature on the volume of a gas at constant pressure, he made an interesting observation. He noted that when he extrapolated the line on a temperature-volume graph to zero volume (V = 0), the line always intersected the *x*-axis at –273°C. This value for temperature is known as 0 on the Kelvin scale. Regardless of the gas tested, Charles also found that for every 1 degree Celsius change, the initial volume was increased or decreased by $\frac{1}{273}$. When the temperature was increased by 273˚C, the volume doubled.

Absolute Zero

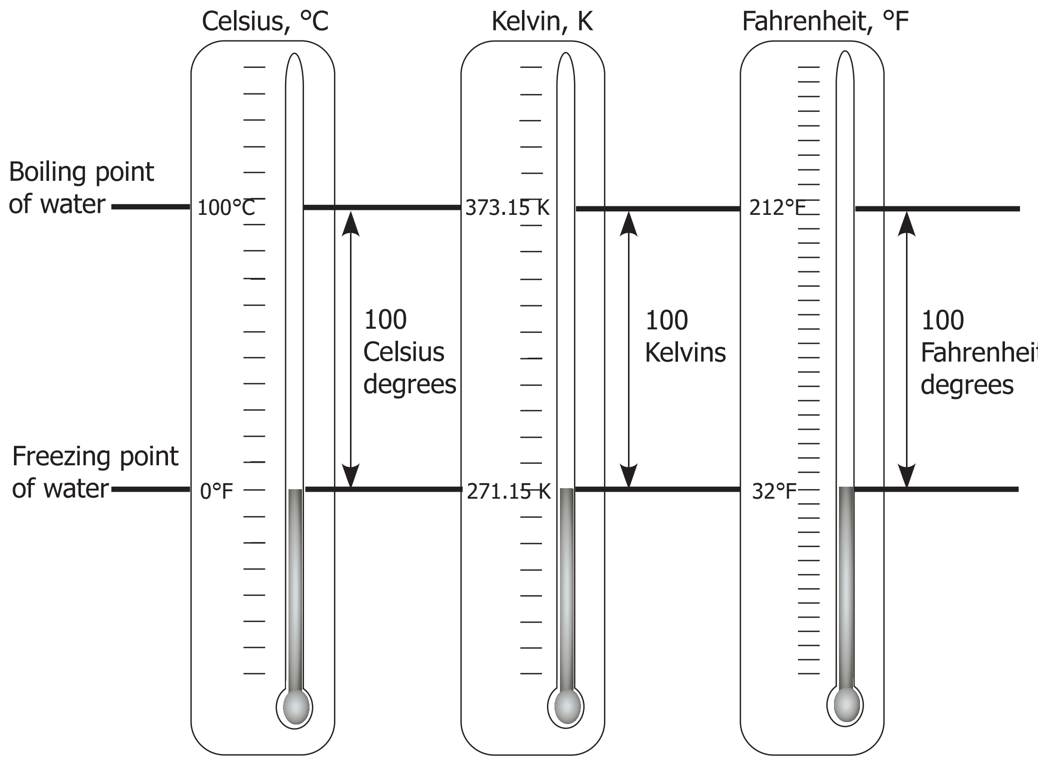
If you look back to Graph 2 in the previous lesson, you will notice the line created by the data should correlate to an x-intercept at about −273°C (the accepted value is −273.15° C).



In 1848, 61 years after Charles' discovery, **William Thomson** (1824-1907) (later known as Lord Kelvin) recognized the significance of the –273 and created the Kelvin scale where –273°C was the lowest temperature possible, which he called absolute zero. Based on this, the *x*-intercept on a graph of volume versus temperature for any gas would always be –273°C 0 Kelvins). Lord Kelvin further reasoned that *at this temperature all molecular motion would cease, the kinetic energy would be zero, and the volume of a gas, hypothetically, would also be zero*. The advantage of this scale is that there are no negative numbers. Here is another version of Graph 2, with temperature measured in Kelvin rather than degrees Celsius.



The diagram below shows a comparison of the three temperature scales:



To convert from Celsius to Kelvin: Add 273

Kelvin = Celsius + 273 ***or*** K = °C + 273

Example 1

Convert −10.0°C to Kelvin.

Solution: K = °C + 273 = −10.0°C + 273 = 263 K

Note: There is no degree sign for temperatures in Kelvin. We say the temperature is "263 Kelvins," no degrees.

To convert from Kelvin to Celsius: Subtract 273

Example 2

Convert 298 K into degrees Celsius.

Solution: °C = K − 273 = 298 − 273 = 25°C

Charles's Law states that the ratio of $\frac{V_1}{T_1}$ is equal to the ratio of $\frac{V_2}{T_2}$, as long as pressure is constant. This *mathematical relationship* is known as the **Charles' Law equation**, and can be expressed as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Since the ratio of the variables is always constant, this implies the relationship between them is direct and that the graph will always show a straight line. Recall that in a direct relationship, the ratio of the two variable quantities is constant. This means that we can compare any two sets of data for volume and temperature and calculate the ratio to be the same.

Using the second version of Graph 2 above, you can prove this to yourself by comparing two sets of data. For example, the third point on the graph corresponds to the following data:

$$\frac{V_1}{T_1} = \frac{200 \text{ mL}}{100 \text{ K}} = 2 \text{ mL/K}$$

The fifth point on the graph demonstrates the same ratio:

$$\frac{V_2}{T_2} = \frac{500 \text{ mL}}{250 \text{ K}} = 2 \text{ mL/K}$$

According to Charles' Law, there is a direct relationship between the temperature and the volume of a gas at a constant pressure. That is, if the Kelvin temperature of a gas were doubled, the volume of the gas would double as well.

Solving Problems Using Charles' Law

We solve problems using Charles' Law in much the same manner that we solved the problems using Boyle's Law:

1. *Analyze the data,* listing both the known and unknown variables.
2. *Set up the two possible ratios.*
3. *Evaluate* whether your answer is logical.

It is important to note that the temperature must ***always*** be in Kelvin. If temperature is in Celsius, the ratio of the variables will not be direct and you will not see a straight line on your graph. Therefore, before you start to solve any problem, you may first need to convert temperature from degrees Celsius to Kelvin.

The example below illustrates why we must use the Kelvin temperature.

Example 3

What is the new volume of a gas if 100 mL of a gas at 25°C is cooled to −25°C?

Step 1: *Analyze the data.*

V_1 = 100 mL

V_2 = ?

T_1 = 25ºC

T_2 = –25ºC

Step 2: *Set up the two possible ratios.* Since the temperature is decreasing from 25°C to −25°C, the volume must also decrease.

We can use two ratios:

$$\frac{25°\text{C}}{-25°\text{C}} \text{ or } \frac{-25°\text{C}}{25°\text{C}}$$

You may notice two problems with these ratios. First of all, neither ratio will result in a lower volume since 25 ÷ 25 = 1. In addition, the resulting volume will have a negative value, and you cannot have a negative volume! To avoid these problems, *you must FIRST convert all temperatures to Kelvin*.

25°C + 273 = 298 K

−25°C + 273 = 248 K

There are two possible ratios:

$$\frac{298\text{ K}}{248\text{ K}} \text{ or } \frac{248\text{ K}}{298\text{ K}}$$

Step 3: *Evaluate your answer*. Since the volume of gas must decrease as temperature decreases, you must use the second ratio.

$$\text{new volume} = 100\text{ mL}\left(\frac{248\ \cancel{\text{K}}}{298\ \cancel{\text{K}}}\right) = 83.2\text{ mL}$$

The new volume (V_2) is about 83.2 mL.

Example 4

If the volume of a gas at −74.0°C is doubled to 48.0 L, calculate the final temperature in degrees Celsius.

Solution:

Step 1: *Analyze the data.* If the volume doubled to 48.0 L, its initial volume must have been half of 48.0 L, or 24.0 L.

V_1 = 24.0 L

V_2 = 48.0 L

T_1 = –74.0°C

T_2 = ?

Step 2: *Set up the two possible ratios.*

The first step is to convert the Celsius temperature to Kelvin.

$$-74.0°\text{C} + 273 = 199 \text{ K}$$

The volume increases, so according to Charles' Law the temperature must have increased. There are two possible ratios:

$$\frac{24.0 \text{ L}}{48.0 \text{ L}} \text{ or } \frac{48.0 \text{ L}}{24.0 \text{ L}}$$

Step 3: *Evaluate your answer.* Since the temperature increases, you must use the second ratio.

$$\text{new temperature} = 199 \text{ K} \left(\frac{48.0 \cancel{\text{L}}}{24.0 \cancel{\text{L}}} \right) = 398 \text{ K}$$

Finally, you must convert the Kelvin temperature back to Celsius.

$$\text{K} - 273 = 398 - 273 = 125°\text{C}$$

Learning Activity 2.6

Working with Charles' Law

1. Calculate the final temperature in degrees Celsius when the volume of 1.00 L of a gas, at 25.0°C, is doubled to 2.00 L.
2. Calculate the volume of 100.0 mL of a gas if its temperature is doubled from 20.0°C to 40.0°C.
3. You get a balloon at the circus with a volume of 2.50 L at room temperature (25.0°C). You then step outside on a cold winter day (–25.0°C). If the pressure remains constant, what is the balloon's new volume?
4. 20.0 mL of a gas at 285 K increases in volume to 32.0 mL. What is the new temperature of the gas?
5. Find the new volume if the following changes occur at a constant pressure:
   a) 225 mL of oxygen at 273 K is warmed to 301 K.
   b) 3.0 L of nitrogen is cooled from 90.0°C to −45.0°C.
6. Would it be possible for a sample of gas to have a volume of zero? Explain.

Lesson Summary

In this lesson, you learned that when the pressure and amount of a gas are unchanged, the ratio of the volume of the gas to the temperature of the gas is constant. This relationship is known as Charles' Law. In the next lesson, you will investigate the relationship between gas pressure and temperature.

NOTES

Assignment 2.5

Solving Problems with Charles' Law *(10 marks)*

For each question that involves a calculation, one mark is for showing your work. Each answer is worth one mark and must reflect the correct unit and number of significant figures.

1. Find the new temperature in degrees Celsius when the following changes occur at a constant pressure:

   a) 5.00 L of air at 45.0°C expands to 22.0 L. *(2 marks)*

   b) 100.0 mL of helium at 301 K compresses to 80.0 mL. *(2 marks)*

continued

2. A sample of gas occupies 6.65 L at 325ºC. What will the new volume be at 25.0ºC if the pressure of the gas does not change? *(2 marks)*

3. 4.00 L of air at −50.0ºC is warmed to 90.0ºC. What is the new volume of gas if the pressure does not change? *(2 marks)*

continued

Assignment 2.5: Solving Problems with Charles' Law (continued)

4. While solving Charles' Law problems, remember that the new volume of a gas is equal to its original volume times a quotient. Is this ratio smaller than one or larger than one if

   a) the gas is cooled? *(1 mark)*

   ____

   ____

   ____

   b) the gas is warmed? *(1 mark)*

   ____

   ____

   ____

NOTES

Lesson 8: Gas Pressure and Temperature, Part A (1.5 hours)

Lesson Focus

SLO C11-2-07A: Experiment to develop the relationship between the pressure and temperature of a gas using visual and graphical representations.
Include: historical contributions of Joseph Louis Gay-Lussac

Lesson Introduction

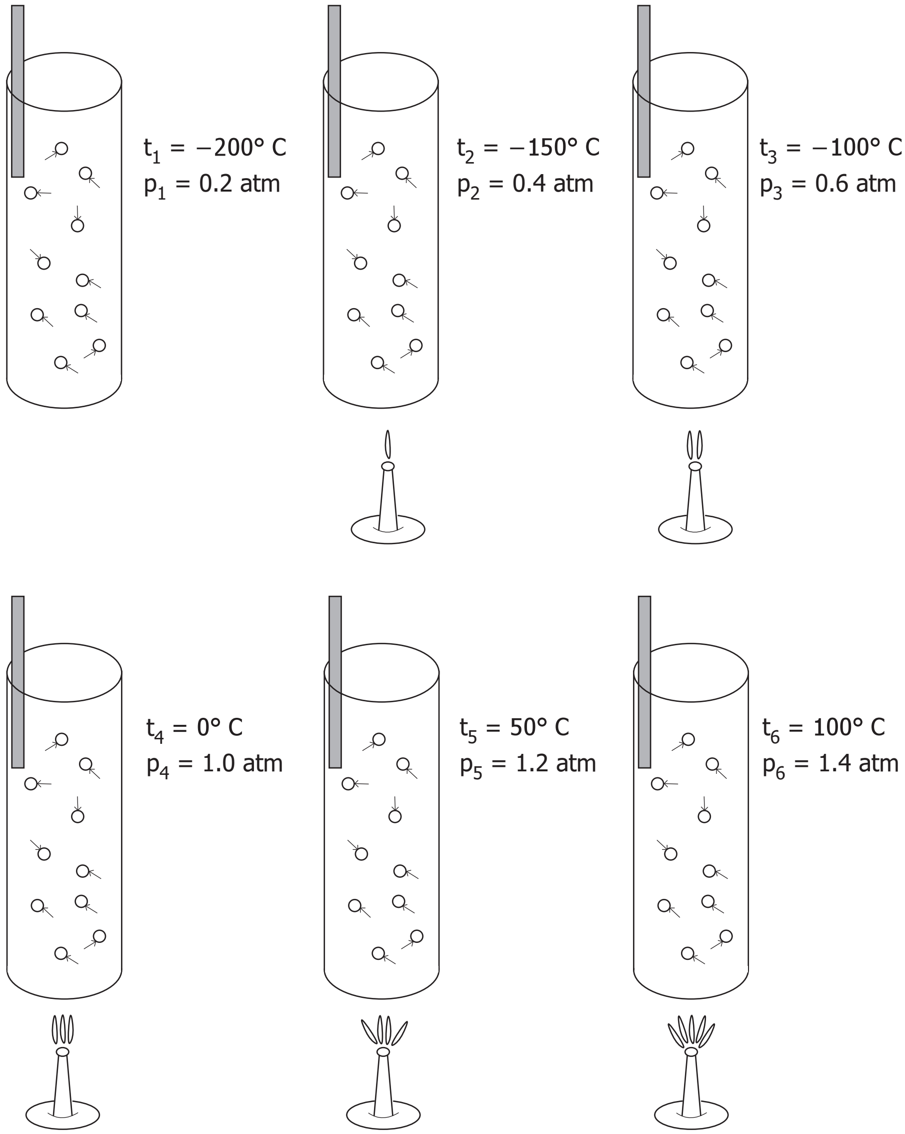
Why is it recommended to check the pressure in your tires before driving your car? Have you ever noticed that the warning label on an aerosol can advises not to store the can above a certain temperature? Both of these scenarios can be explained as a result of the work of Joseph Gay-Lussac, who studied the relationship between the pressure and temperature of a contained volume of gas. In the case of tire pressure, as the air inside the tire warms while driving, the pressure inside the tire increases. Thus, the most accurate reading of tire inflation is before you drive anywhere. Likewise, pressure will build inside of aerosol cans that are heated, increasing the danger of an explosion.

Joseph Gay-Lussac

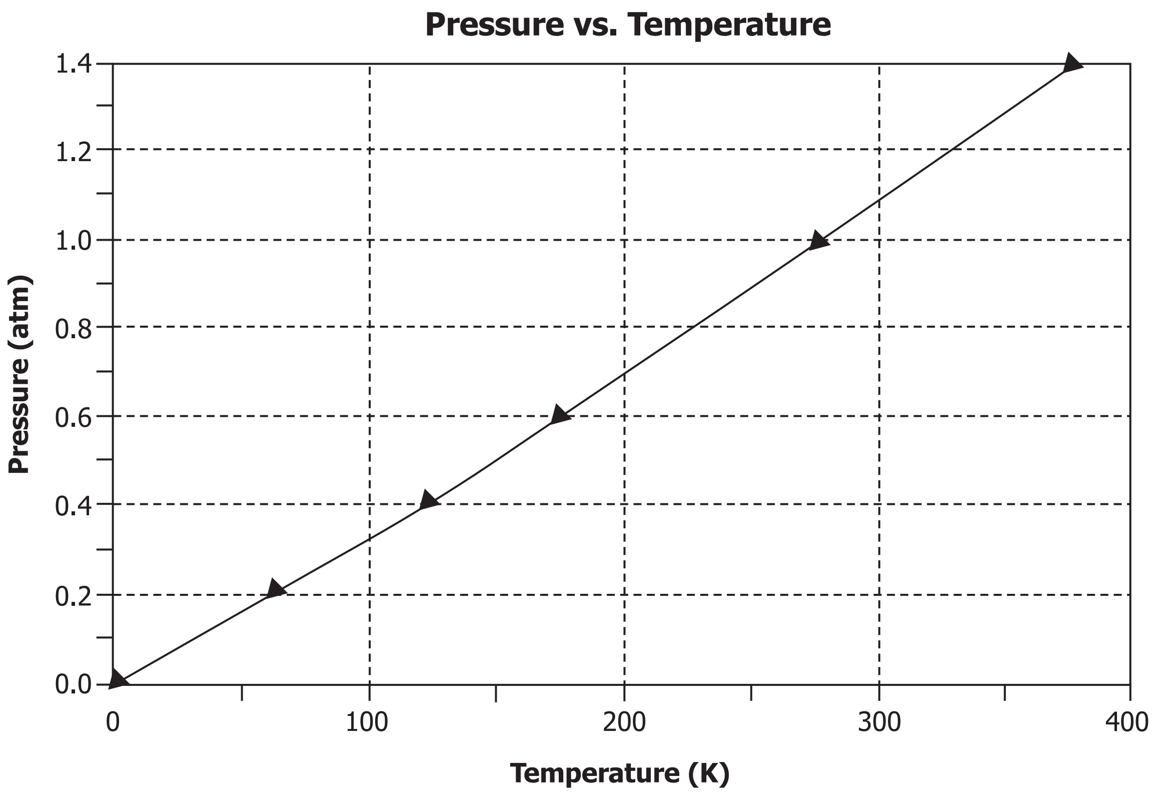
Robert Boyle discovered that pressure and volume were inversely related (as one variable increases, the other decreases), while Jacques Charles discovered that pressure and volume were directly related (both variables increase at the same time). Joseph Gay-Lussac (1778-1850) carried on Charles' work and discovered the relationship between temperature and gas pressure.

Gay-Lussac became well known for his dedication to meticulous experimentation. Evidence of this was his desire to remove all water vapour from the gases in his experiments. As a result, he observed more consistent and reproducible results than Charles had. With respect to the thermal expansion of gases, Gay-Lussac's precise measurements were used by Lord Kelvin to determine the value of absolute zero.

As an extension to his studies on the expansion of gases, Gay-Lussac determined that if the volume and the amount of a gas are held constant, increasing the temperature of a gas will increase the pressure—another direct relationship. This became known as **Gay-Lussac's Law**.



If you graphed the pressure and temperature data represented by the previous cylinders, you would find that there is a linear relationship between the variables. This relationship is demonstrated below in the Pressure vs. Temperature graph. We will refer to this as Graph 3.



Using data from Graph 3 above, you can determine that the ratio of pressure and temperature remains constant as long as the volume of the gas remains unchanged. For example, the third point on the graph corresponds to the following:

$$\frac{P_1}{T_1} = \frac{0.4 \text{ atm}}{125 \text{ K}} = 0.003 \text{ atm/K}$$

The fourth set of data on the graph corresponds to the following:

$$\frac{P_2}{T_2} = \frac{0.6 \text{ atm}}{175 \text{ K}} = 0.003 \text{ atm/K}$$

Explaining Gay-Lussac's Law

If we examine the temperature-pressure relationship of a gas with respect to the kinetic molecular theory, an increase in the temperature of a gas increases the kinetic energy of the gas and thus the speed of the particles and the frequency and force of collisions with the sides of the container. We would then expect that an increase in temperature should cause an increase in the pressure of a gas at constant volume because the increased frequency of collisions would increase the force applied to the sides of the container. This is why it is dangerous to heat an aerosol can. If the temperature is increased sufficiently, the increase in pressure may cause the can to explode.

The popular "Dancing Penny" demonstration takes advantage of Gay-Lussac's Law. It begins when a wet penny is placed in top of the mouth of a glass bottle. Heat from the demonstrator's hand causes the air pressure of the gas inside the bottle to increase until the penny pops up slightly. This phenomenon can continue for several minutes. If you have access to the Internet and have never seen this demonstration, you can access a video at www.metacafe.com/watch/1001831/simple_science_ the_dancing_ penny/.

Another example of Gay-Lussac's Law at work involves specific tire pressures. All cars—race cars in particular—require specific tire pressures for ideal handling. For example, drag racing slicks are inflated to only 12 or 15 p.s.i.. When cold tires are placed on the dragster, they are filled to a lower than ideal pressure. As the dragster races, the friction will heat up the tires and the pressure will be raised to the ideal pressure. Sometimes race car tires are even pre-warmed to reduce the amount of time that it takes the tires to get up to the temperature that will produce the ideal pressure.

We don't all drive race cars, but the recommended pressure found on the side of your tires is meant to be cold tire pressure. Most passenger vehicles require a tire pressure of 32 to 35 p.s.i., while full-size trucks and sport utility vehicles require about 40 p.s.i..

Lesson Summary

Understanding the relationship between the temperature and pressure of a gas can help you understand why heating an aerosol container can be dangerous and why tire pressure should be measured when the tire is cold. In the next lesson, you will continue working with Gay-Lussac's Law to solve problems involving changes in gas pressure and temperature.

Assignment 2.6

Investigating the Temperature-Pressure Relationship *(10 marks)*

Use the Pressure vs. Temperature graph (Graph 3) from the lesson to help you answer the following questions.

1. What do you notice about the temperature at zero pressure on the graph? What is the significance of this temperature? *(2 marks)*

2. What do you notice about the pressure at zero degrees *Celsius* on the graph? *(1 mark)*

3. Is the relationship between gas pressure and temperature inverse or direct? How do you know by looking at the graph? *(2 marks)*

4. What happens to gas pressure as temperature increases? *(1 mark)*

continued

Assignment 2.6: Investigating the Temperature-Pressure Relationship (continued)

5. Complete the following statement with the correct terms. *(2 marks)*

   Gas particles increase the frequency and force of their collisions as heat ________________ . This increased kinetic energy also causes increased gas ________________ at a constant volume.

6. Describe what happens to an inflated balloon when it is taken outside on a cold day. How does this illustrate the findings of Gay-Lussac? *(2 marks)*

   ____

   ____

   ____

Lesson 9: Gas Pressure and Temperature, Part B (1.5 hours)

Lesson Focus

SLO C11-2-07B: Experiment to develop the relationship between the pressure and temperature of a gas using numeric representation.
Include: historical contributions of Joseph Louis Gay-Lussac

Lesson Introduction

Gay-Lussac's Law helps explain why food cooks faster in a pressure cooker, where the volume of food is constant but the pressure increases rapidly. In the previous lesson, you observed that as gas pressure increases, so does the temperature. In this lesson, you will use mathematics to help you solve problems using this direct relationship.

Solving Problems Using Gay-Lussac's Law

Gay-Lussac's Law states that the pressure of a gas is directly proportional to the Kelvin temperature as long as the volume of gas remains constant. Since this law involves direct proportions, the ratios $\frac{P_1}{T_1}$ and $\frac{P_2}{T_2}$ are equal at constant volume. This *mathematical relationship* is known as the **Gay-Lussac's Law equation**, which can be expressed as:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

As with Boyle's Law and Charles' Law, when solving problems, we must first determine whether the change identified will increase or decrease the original value. Then we can determine which ratio will produce the appropriate change.

As with Charles' Law, temperature values must be in Kelvin.

Example 1

If a 12.0 L sample of gas is found to have a pressure of 101.3 kPa at 0.0˚C, calculate the new pressure at 128˚C if the volume is held constant.

Step 1: *Analyze the data.*

P_1 = 101.3 kPa

P_2 = ?

T_1 = 0.0ºC

T_2 = 128ºC

Step 2: *Set up the two possible ratios.*

First, you must convert temperature values to Kelvin:

T_1 = 0.0°C + 273 = 273 K

T_2 = 128°C + 273 = 401 K

Temperature increases, so pressure must increase. There are two possible ratios:

$$\frac{273\text{ K}}{401\text{ K}} \text{ or } \frac{401\text{ K}}{273\text{ K}}$$

Step 3: *Evaluate your answer.* Only the second ratio will increase the pressure, so:

$$\text{new pressure} = 101.3\text{ kPa}\left(\frac{401\ \cancel{\text{K}}}{273\ \cancel{\text{K}}}\right) = 149\text{ kPa}$$

Example 2

A can of hairspray is at a pressure of 103 kPa at 22ºC. Someone throws the can into a fire, and the can reaches a pressure of 432 kPa. What will be the new temperature of the can at this pressure?

Step 1: *Analyze the data.*

P_1 = 103 kPa

P_2 = 432 kPa

T_1 = 22ºC

T_2 = ?

Step 2: *Set up the two possible ratios.*

First, you must convert temperature values to Kelvin:

T_1 = 22°C + 273 = 295 K

There are two possible ratios:

$$\frac{103\ \text{kPa}}{432\ \text{kPa}} \text{ or } \frac{432\ \text{kPa}}{103\ \text{kPa}}$$

Step 3: *Evaluate your answer.* Since the pressure increases, the temperature must also increase. The second ratio will increase the temperature.

$$\text{new temperature} = 295\ \text{K}\left(\frac{432\ \text{kPa}}{103\ \text{kPa}}\right) = 1237.3\ \text{K} = 1240\ \text{K}$$

Convert Kelvin to Celsius:

1240 K – 273 = 967°C ≈ 970°C

Learning Activity 2.7

Gay-Lussac's Law

1. A gas in a rigid container has a pressure of 1.00 atm at 25.0°C. What is the pressure of the gas at 50.0°C?
2. A gas in a rigid container has a pressure of 695 torr at 19.0°C. What is the pressure of the gas at −20.0°C?
3. The water vapour in a pressure cooker is at standard pressure (101.3 kPa) at 30.0°C. What is the temperature, in degrees Celsius, when the pressure is 151 kPa?
4. The air pressure in a scuba tank is 175 atm at room temperature, 22.0°C. What is the pressure in the tank while diving in water at a temperature of 6.0°C?
5. You fill your tires on a cool morning (10.0°C) to a pressure of 205 kPa. If you take a long driving trip where the tire temperature reaches 75.0°C, what is the pressure in the tire while you are driving?

Lesson Summary

In this lesson, you practised solving problems involving the direct relationship between temperature (which must always be converted to Kelvin) and pressure. In the next lesson, you will solve problems involving gas pressure, temperature, and volume.

Assignment 2.7

Problem Solving with Gay-Lussac's Law *(10 marks)*

Show all of your calculations, including temperature conversions. Be sure to include appropriate units for your answers.

1. While solving Gay-Lussac's Law problems, remember that the new pressure of a gas is equal to its original pressure times a ratio. Is this ratio larger or smaller than one if

   a) the gas is cooled? *(1 mark)*

   ____

   ____

   ____

   b) the gas is warmed? *(1 mark)*

   ____

   ____

   ____

2. Instead of filling the tires on your car when they are cold, you fill them after driving for a period of time. You stop at a gas station and fill your tires to 205 kPa when the tire temperature is 45.0°C. A cold snap hits and the temperature dips to −30.0°C. What is the pressure in the tires? *(4 marks)*

continued

Assignment 2.7: Problem Solving with Gay-Lussac's Law (continued)

3. An air compressor is filled to a pressure of 3.20 atm at 26.0°C. If the air is suddenly released from the compressor and the pressure drops to 2.60 atm, what is the temperature, in degrees Celsius, of the air remaining in the compressor? *(4 marks)*

Lesson 10: The Combined Gas Law (2 hours)

Lesson Focus

SLO C11-2-08A: Solve quantitative problems involving the relationships among the pressure, temperature, and volume of a gas using dimensional analysis.
Include: symbolic relationships

Lesson Introduction

In the previous lessons, you have investigated only two variables relating to gases at a time, such as pressure and volume. In this lesson, you will go one step further and work with a single expression that combines Boyle's, Charles', and Gay-Lussac's laws.

The Combined Gas Law

Up until now you have solved problems where only one variable is unknown. In each of Boyle's, Charles', and Gay-Lussac's laws, there were only two conditions to analyze, such as the volume and temperature of a contained gas. In this lesson, we will solve problems involving changes in volume, pressure, and temperature. This is commonly known as the combined gas law. **The Combined Gas Law** describes the relationship among the pressure, volume, and temperature of a contained gas.

There are many situations where a sample of gas is subjected to changes in both pressure and temperature at the same time. As a result, the volume of the gas can change. For example, as a weather balloon rises into the atmosphere to collect meteorological data, the air temperature drops and the volume of the gas decreases. At the same time, the atmospheric pressure acting on the balloon decreases, which allows the gas to expand. As both statements are conflicting, you can see that there is no single gas law that you have learned that would allow you to determine the final volume of the gas in the weather balloon. Using the Combined Gas Law, though, you would be able to determine which variable has the greatest effect on the volume of the gas.

Solving Problems with Changing Temperature, Volume, and/or Pressure

Now that you have solved problems using each gas law individually, you can use them together. Just remember the relationships:

Boyle's Law: Pressure and volume are indirectly related ($V \alpha 1/P$)
Charles' Law: Volume and temperature are directly related ($V \alpha T$)
Gay-Lussac's Law: Pressure and temperature are directly related ($P \alpha T$)

Remember that for Boyles' Law, temperature is constant. When using both Boyle's and Charles' laws, temperature must be in Kelvin.

Example 1

If a gas occupies a volume of 25.0 L at 25.0°C and 1.25 atm, calculate the volume at 128°C and 0.750 atm.

Step 1: *Analyze the data.*

You are given:

25.0 L at 25.0°C and 1.25 atm

You need to find:

Volume at 128°C and 0.750 atm

Step 2: *Set up the two possible ratios.*

In order to solve problems where more than one variable changes, you must treat each change separately, but in the same way you did in the previous lessons. In other words, you can use each gas law separately as a two-step process to solve this problem.

Step A: In the first relationship, temperature goes from 25.0°C to 128°C.

Convert from Celsius temperatures to Kelvin:

25.0°C + 273 = 298 K

128°C + 273 = 401 K

There are two possible ratios for temperature:

$$\frac{298\text{ K}}{401\text{ K}} \text{ or } \frac{401\text{ K}}{298\text{ K}}$$

Step B: In the second relationship, the pressure changes from 1.25 to 0.750 atm. The two ratios for pressure are:

$$\frac{1.25 \text{ atm}}{0.750 \text{ atm}} \text{ or } \frac{0.750 \text{ atm}}{1.25 \text{ atm}}$$

Step 3: *Evaluate your answer.*

According to Charles' Law, increasing temperature causes the volume to increase. The second ratio from Step A will increase volume:

$$\text{new volume} = 25.0 \text{ L} \times \left(\frac{401 \cancel{\text{K}}}{298 \cancel{\text{K}}}\right) \times ?$$

According to Boyle's Law, if pressure decreases, volume increases. This is reflected by the first ratio from Step B, which can then be multiplied by the appropriate ratio from Step A:

$$\text{new volume} = 25.0 \text{ L} \times \left(\frac{401 \cancel{\text{K}}}{298 \cancel{\text{K}}}\right) \times \left(\frac{1.25 \cancel{\text{atm}}}{0.750 \cancel{\text{atm}}}\right) = 56.1 \text{ L}$$

Example 2

If a gas has a volume of 125 L at 325 kPa and 58.0°C, calculate the temperature in Celsius to produce a volume of 22.4 L at 101.3 kPa.

Step 1: *Analyze the data.*

You are given:

125 L at 325 kPa and 58.0°C

You need to find:

Temperature at 22.4 L at 101.3 kPa

Step 2: *Set up the two possible ratios.*

Step A: You must first convert from Celsius temperatures to Kelvin:

58.0°C + 273 = 331 K

In the first relationship, volume goes from 125 L to 22.4 L. There are two possible ratios for volume:

$$\frac{22.4 \text{ L}}{125 \text{ L}} \text{ or } \frac{125 \text{ L}}{22.4 \text{ L}}$$

Step B: The pressure changes from 325 kPa to 101.3 kPa in the second relationship. There are two possible ratios for pressure:

$$\frac{325 \text{ kPa}}{101.3 \text{ kPa}} \text{ or } \frac{101.3 \text{ kPa}}{325 \text{ kPa}}$$

Step 3: *Evaluate your answer.*

According to Charles' Law, if volume is decreased, temperature should decrease as well. The first ratio from Step A decreases temperature:

$$\text{new temperature} = 331 \text{ K} \times \left(\frac{22.4 \ \cancel{\text{L}}}{125 \ \cancel{\text{L}}}\right) = ?$$

Guy-Lussac's Law says that if pressure decreases, temperature should also decrease. You should use the second ratio from Step B and multiply it by the appropriate ratio from Step A:

$$\text{new temperature} = 331 \text{ K} \times \left(\frac{22.4 \ \cancel{\text{L}}}{125 \ \cancel{\text{L}}}\right) \times \left(\frac{101.3 \ \cancel{\text{kPa}}}{325 \ \cancel{\text{kPa}}}\right) = 18.5 \text{ K}$$

Finally, you must convert your Kelvin temperature to Celsius:

$$18.5 \text{ K} - 273 = -254°\text{C}$$

Using Symbolic Relationships

While investigating each gas law, you have focused on ratios and relationships when solving problems. At the same time, you were introduced to the symbolic relationships (mathematical equations) used to describe each gas law. Let's review what you have seen so far.

Boyle's Law

If you start with a sample of gas with a certain volume (V_1) and pressure (P_1), then change the volume to a new value (V_2), the pressure will change to a new value (P_2). The products of both pressure-volume values will be equal since they are from the same sample at the same temperature. The indirect relationship between pressure and volume can be expressed mathematically as

$$P_1V_1 = P_2V_2$$

Charles' Law

If you start with a volume of gas (V_1) at a specific temperature (T_1), and then change the temperature to T_2, its volume will change to V_2. This means that for a sample of gas with a constant pressure, $V \div T$ will result in a constant value. The direct relationship between volume and temperature can be expressed mathematically as

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Gay-Lussac's Law

Like volume and temperature, temperature and pressure are directly related. Since the relationship results in a constant value, a change in pressure or temperature of a sample of gas at a constant volume can be expressed mathematically as

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

The Combined Gas Law

These three relationships can be combined into one mathematical relationship:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This *mathematical relationship* is known as the **Combined Gas Law equation**. You can use this equation to solve problems similar to the first two examples in this lesson.

Example 3

Salvage divers often use lift bags to lift objects to the surface. Divers are required to make a pre-dive calculation of the forces involved, to ensure the safety of the divers during the recovery. In one instance, a lift bag contains 145 L of air at the bottom of a lake, where there is a temperature of 5.20°C and a pressure of 6.00 atm. As the bag is released, it ascends to the surface where the pressure is 1.00 atm and the temperature is 16.0°C. Calculate the volume the gas would occupy at the surface of the lake. If the maximum volume of the lift bag is 750 L, will the bag burst at the surface?

Step 1: *Analyze the data.* Convert temperature(s) to Kelvin.

$P_1 = 6.00$ atm

$P_2 = 1.00$ atm

$T_1 = 5.20°C + 273 = 278.2$ K

$T_2 = 16.0°C + 273 = 289$ K

$V_1 = 145$L

$V_2 = ?$

Step 2: *Solve using the Combined Gas Law* by substituting known values into the equation and solving for the unknown value.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(6.00 \text{ atm})(145 \text{ L})}{278.2 \text{ K}} = \frac{(1.00 \text{ atm})V_2}{289 \text{ K}}$$

Rearrange to solve for V_2:

$$V_2 = \frac{(6.00 \text{ atm})(145 \text{ L})(289 \text{ K})}{(278.2 \text{ K})(1.00 \text{ atm})}$$

$$V_2 = 904 \text{ L}$$

Step 3: *Evaluate your answer.*

Boyle's Law says that if pressure decreases, volume should increase, while Charles' Law states that volume increases when temperature increases. The volume did, in fact, increase from V_1 to V_2. The new volume of 904 L will, in fact, cause the bag to burst, since it exceeds the 750 L limit of the lift bag.

Learning Activity 2.8

The Combined Gas Law

Solve each of the following problems using the ratio method *and* the Combined Gas Law. Include appropriate units for your final answer.

1. In a laboratory experiment, 85.3 mL of a gas are collected at 24.0°C and 733 mmHg pressure. Find the volume at 760 mmHg pressure and 0.0ºC.
2. A flexible container has a maximum volume of 15.0 L. If the container is filled to 8.00 L at a pressure of 1.80 atm and a temperature of 17.0°C, at what temperature will the container have a maximum volume at a pressure of 2.90 atm?
3. If you have 17.0 L of gas at a temperature of 28.0°C and a pressure of 88.9 atm, what will be the pressure of the gas if you raise the temperature to 68.0°C and decrease the volume to 12.0 L?
4. A gas that has a volume of 28.0 L, a temperature of 45.0° C, and an unknown pressure has its volume increased to 34.0 L and its temperature decreased to 35.0°C. If the pressure after the change is 2.00 atm, what was the original pressure of the gas?

Lesson Summary

In this lesson, you learned how to solve problems where only the amount of gas was constant. You practiced the ratio method, where the three gas laws can be used separately to solve a problem. You also learned about the Combined Gas Law and how to use this symbolic equation to help you solve problems.

NOTES

Assignment 2.8

Problem Solving with the Combined Gas Law *(28 marks)*

Solve each of the following problems using the ratio approach OR the combined gas law as indicated in the problem. Include appropriate units for your final answer. Each problem has a value of 7 marks. For each solution, marks will be allotted for the following:

- List givens *(1 mark)*
- Convert temperature(s) *(2 marks)*
- Write a prediction of the outcome (Example“Since the volume of gas increases, so will the gas pressure.”) *(1 mark)*
- Set up ratios OR equation *(1 mark)*
- Plug in correct values *(1 mark)*
- Provide a final answer with correct units and significant figures *(1 mark)*

1. Use the <u>*ratio method*</u> to solve the following problem: A sample of argon gas has a volume of 205 mL when its temperature is –44.0°C and its pressure is 712 mmHg. What would be the volume of the argon at 755 mmHg and –15.0°C?

continued

Assignment 2.8: Problem Solving with the Combined Gas Law (continued)

2. Use the *ratio method* to solve the following problem: A student collects a 325 mL sample of hydrogen gas at 36.0°C and 103.2 kPa. What volume would the hydrogen occupy at 91.9 kPa and 18.0°C?

continued

Assignment 2.8: Problem Solving with the Combined Gas Law (continued)

3. Use the <u>*Combined Gas Law equation*</u> to solve the following problem: A toy balloon has an internal pressure of 1.05 atm and a volume of 5.00 L at 20.0°C. The balloon is released and reaches the upper atmosphere where the volume of the balloon becomes 21.0 L and the pressure is 0.210 atm. What is the temperature at this altitude?

continued

Assignment 2.8: Problem Solving with the Combined Gas Law (continued)

4. Use the *Combined Gas Law equation* to solve the following problem: A gas collected on a day when the pressure was 101 kPa and the temperature was 8.00°C has a volume of 942 mL. If the volume on another day changed to 837 mL when the temperature was 33.0°C, what was the pressure on that day?

Lesson 11: Application of the Gas Laws (1 hour)

Lesson Focus

SLO C11-2-09: Identify various industrial, environmental, and recreational applications of gases.
Examples: self-contained underwater breathing apparatus (scuba), anaesthetics, air bags, acetylene welding, propane appliances, hyperbaric chambers...

Lesson Introduction

There are many uses of gases in our daily lives. How would your life be different without gases? Read on to find out more about a few of the fascinating ways gases are used in industry, recreation, and the environment.

Gases in Industry

How can industry lend a hand in providing year-round transportation to remote and inaccessible regions? For example, many communities in Northern Manitoba require shipments of goods year-round, but are isolated when winter roads over lakes are dangerous and melt earlier than usual. One solution to this dilemma is the airship, a technology that could move shipments of construction materials, food products and other heavy materials to remote areas.

You may have heard of the zeppelin, a rigid airship designed by the Germans. Early versions of the zeppelin were used as passenger transport and then later as bombers and scouts in the First World War. Smaller, remote-controlled versions of airships are used today at sporting events, where a camera hovers overhead to capture plays in progress.

These airships are often called “lighter-than-air craft” since they literally float on air. To achieve lift in the air, the body of the airship must be filled with a gas that is lighter than air. The safest choices are usually hydrogen and helium gases, or a combination of the two, which are generally neither toxic nor corrosive. There are also thermal airships that heat the air in the cavity of the ship to achieve lift, similar to a hot air balloon.

Modern airships use “dynamic helium volume.” At sea level altitude, helium only takes up a small part of the hull, while the rest of the volume is filled with air. As the airship ascends, and the atmospheric pressure decreases, the helium expands. As a result, the air is pushed out of the hull through a valve underneath the airship. This allows airships to reach altitudes of over 30,000 m with balanced inner and outer pressure, without exploding due to an overloaded inner pressure.

The development of airship technology could provide roadless trucking for communities all over the world, improving the standard of living. This technology is perhaps a more environmentally friendly solution to the paving of highways, and is certainly a more cost-effective means of moving goods.

Gases and Recreation

How long can you stay underwater after taking a deep breath of air? Likely not long enough to explore a coral reef or a sunken vessel. In those situations, you would need a tank of compressed air. The air in your tank would be a mixture of gases, the composition of which would vary based on the depth and length of your dive.

As a diver descends, there is an increased pressure on the lungs of the diver. More internal gas pressure is then required to keep the lungs expanded. This increased gas pressure causes oxygen and nitrogen from the air mixture to be dissolved in the diver’s blood. Dissolved oxygen is not problematic, since body cells remove it from the blood. But what about nitrogen, which is not metabolized by the human body? The extra nitrogen that is inhaled has nowhere to go but into the blood and tissues, where it stays in the gas phase ("dissolved") at the higher pressure. To avoid this dangerous situation, the regulator, which is placed in the mouth, adjusts the pressure of the gas mixture to keep the pressure inside the lungs equal to the pressure outside the lungs.

As the diver ascends following the dive, the pressure decreases and the dissolved nitrogen gas is released from the blood. Because dissolved nitrogen can only be excreted through the lungs at a controlled rate, bubbles of the released gas can cause problems by blocking small blood vessels and limiting oxygen uptake by the cells. If this happens, a diver may experience decompression sickness, which causes severe joint pain, vomiting, and dizziness. In order to avoid **decompression sickness**, divers have to pay careful attention to the length and frequency of their dives, and must ascend to the surface at an appropriate rate. A speedy ascent may not allow expanding gas bubbles to be excreted safely.

In order to understand this concept better, you can check out the video on diving at www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/gam2s3_2.rm.

Learning Activity 2.9

The Gas Laws at Work

For each of the following applications, identify the related gas law. Use your course notes as well as other resources, if necessary, to help you.

1. A diver swimming at a depth of 10 m under the water observes that the size of his air bubbles increase in volume as they rise to the water's surface.

   Gas Law: ____________________________

2. As a diver swims deeper into the water, added pressure can cause more nitrogen to be absorbed by your blood than normal. If the diver ascends too quickly, the decrease in pressure can cause those dissolved nitrogen bubbles to expand. This can block the flow of blood to critical organs, cause severe joint pain (known as "the bends"), or produce effects similar to alcohol intoxication (called nitrogen narcosis).

   Gas Law: ____________________________

continued

Learning Activity 2.9 (continued)

3. Nitrogen is pumped into race car tires because it has a more consistent rate of expansion and contraction as the tires heat up on the track and cool down after racing. The expansion and contraction of nitrogen gas creates pressure changes in the tire that are more predictable than when a regular air mixture is used.

   Gas Law: ____________________________

4. "Air blasters" are commercial dusting sprays. They contain a large volume of air that has been compressed into a small space. When the nozzle is pressed, some of the gas escapes from the can.

   Gas Law: ____________________________

5. As a plunger is pushed down into a clogged toilet, the volume of air in the plunger becomes less, and the pressure increases. This increase in pressure is transferred through the water in the drain, which breaks apart the clog.

   Gas Law: ____________________________

6. As the air in a hot air balloon is heated, its volume expands and the balloon rises, because the heated air is less dense than the cooler air around it.

   Gas Law: ____________________________

7. On a hot summer day, your friend shakes a cold can of pop before handing it to you. You tap a few times on the lid (to increase the pressure on the can so that more carbon dioxide gas is forced to dissolve into the liquid) and then pop the tab slowly (so that the drop in pressure inside the container is not too sudden).

   Gas Law: ____________________________

8. The plunger on a turkey syringe thermometer pops out when the turkey is done, because the volume of air trapped under the plunger increases as the temperature inside the turkey increases.

   Gas Law: ____________________________

continued

Learning Activity 2.9 (continued)

9. An air mattress that is blown up in a cool basement feels firmer when it is moved into a warmer room upstairs.

   Gas Law: ___________________________

10. Throwing an aerosol can into a fire can cause an explosion.

    Gas Law: ___________________________

11. A typical diver's lungs can hold up to 8 L of air. As a diver descends, the air in the diver's cavities (i.e., lungs, sinuses, and middle ears) decreases in volume while the pressure increases. At a depth of 10 m, the same amount of gas will occupy only half the volume (4 L) that it would at sea level.

    Gas Law: ___________________________

12. Scuba shop owners often fill tanks in a water bath where the temperature is approximately the same as wherever the diving will take place. They do this because a full scuba tank will gain about 5 or 6 p.s.i. of pressure for every one degree rise in temperature. A tank of 3000 p.s.i. that is heated to 3500 p.s.i. could easily explode.

    Gas Law: ___________________________

13. Scuba divers carry a small container on their back (0.39 $ft.^3$ or 28.3 L), but have large volumes of breathable air available (80 $ft.^3$ at sea level, enough to fill a phone booth). This is because the air is stored at high pressures, and then released at ambient pressure (1 atm) using a regulator.

    Gas Law: ___________________________

14. A football inflated inside and then taken outdoors on a winter day shrinks slightly.

    Gas Law: ___________________________

continued

Learning Activity 2.9 (continued)

15. In a combustion engine cylinder of constant volume, a dramatic increase in temperature will increase the pressure that the gas exerts on the piston. The compression of this piston is what generates the power that moves the car.

    Gas Law: ___________________________

16. A slightly overinflated raft explodes after being launched in tropical waters.

    Gas Law: ___________________________

17. The gauge pressure in a steel-belted automobile tire is higher after travelling over hot asphalt.

    Gas Law: ___________________________

18. When deep sea fish are brought to the surface where the pressure is lower, they die because the volume of gas in their bodies increases, causing bladders, cells, and membranes to pop.

    Gas Law: ___________________________

19. In the event of an automobile accident, heat from an ignition source will cause sodium azide, silicon dioxide, and potassium nitrate to produce the 67 L of nitrogen gas that fills an air bag in 0.03 seconds.

    Gas Law: ___________________________

20. To submerge a submarine, pumps allow water into special tanks and puts air under pressure to decrease its volume. This increases the density of the submarine so that it “dives” down into the water. To surface, the submarine decreases pressure to allow the air to expand back to its original volume and expel the water. The density of the sub then becomes less than that of water and it floats up to the surface.

    Gas Law: ___________________________

Lesson Summary

In this lesson, you learned how gases are used in many applications. Their unique properties allow them to be used recreationally in scuba diving and in industry, as in the case of airships.

NOTES

Module 2 Summary

Congratulations! You have reached the end of Module 2.

Submitting Your Assignments

It is now time for you to submit Assignments 2.1 to 2.8 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 2 assignments and organize your material in the following order:

- Module 2 Cover Sheet (found at the end of the course Introduction)
- Assignment 2.1: Air Quality Improvement Research
- Assignment 2.2: Determining Significant Digits
- Assignment 2.3: Solving Problems with Boyle's Law
- Assignment 2.4: Investigating the Temperature-Volume Relationship
- Assignment 2.5: Solving Problems with Charles' Law
- Assignment 2.6: Investigating the Temperature-Pressure Relationship
- Assignment 2.7: Problem Solving with Gay-Lussac's Law
- Assignment 2.8: Problem Solving with the Combined Gas Law

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

NOTES

Grade 11 Chemistry (30S)

Module 2: Gases and the Atmosphere

Learning Activity Answer Keys

Module 2: Gases and the Atmosphere

Learning Activity 2.1: Measuring Pressure

Part A: Complete the following multiple choice questions by circling the best possible response.

1. This scientist is credited with inventing the barometer.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
2. This scientist developed the manometer and the first useful vacuum pump.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
3. This scientist discovered that water could only be raised to a certain height in a column.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
4. This scientist discovered that the pressure of the atmosphere changes according to elevation.
   a) Torricelli
   b) Galilei
   c) Pascal
   d) Huygens
5. This scientist theorized that a sample of any gas at the same temperature and pressure will contain the same number of particles.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton

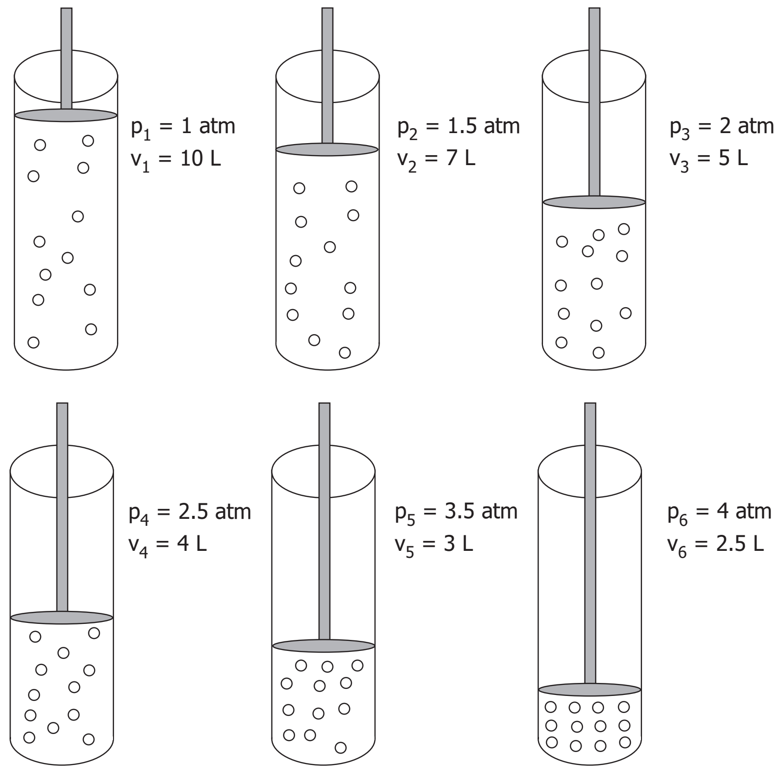
6. This scientist created a very strong vacuum that furthered our understanding of atmospheric pressure.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton
7. This scientist experimented with combining volumes of gases and created a gas law based on his findings.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton
8. This scientist stated that the total pressure of gases in a container is equal to the sum of the pressures of each individual gas.
   a) Von Guericke
   b) Gay-Lussac
   c) Avogadro
   d) Dalton

Part B: Complete the following statements with the correct missing term(s).

1. A measurement unit of pressure is named after ***Blaise Pascal***.
2. A device used to measure air pressure is called a ***barometer***.
3. Atmospheric pressure ***decreases*** as elevation increases.
4. In fair weather and at sea level, the atmospheric pressure is sufficient to support a column of mercury about ***760 mm or 76 cm*** high.
5. How many atmospheres of pressure are required to support 760 mm of mercury in a barometer? ***One***
6. The SI unit for pressure is the ***Pascal***, which represents a small amount of pressure.
7. Normal atmospheric pressure is ***101.3*** kPa.
8. The ***millibar or kilopascal*** is the unit of pressure commonly used in meteorology.
9. ***Pressure*** is defined as force per unit area.
10. Gas pressure is usually measured with an instrument called a ***manometer***.

Learning Activity 2.2: Investigating the Pressure-Volume Relationship

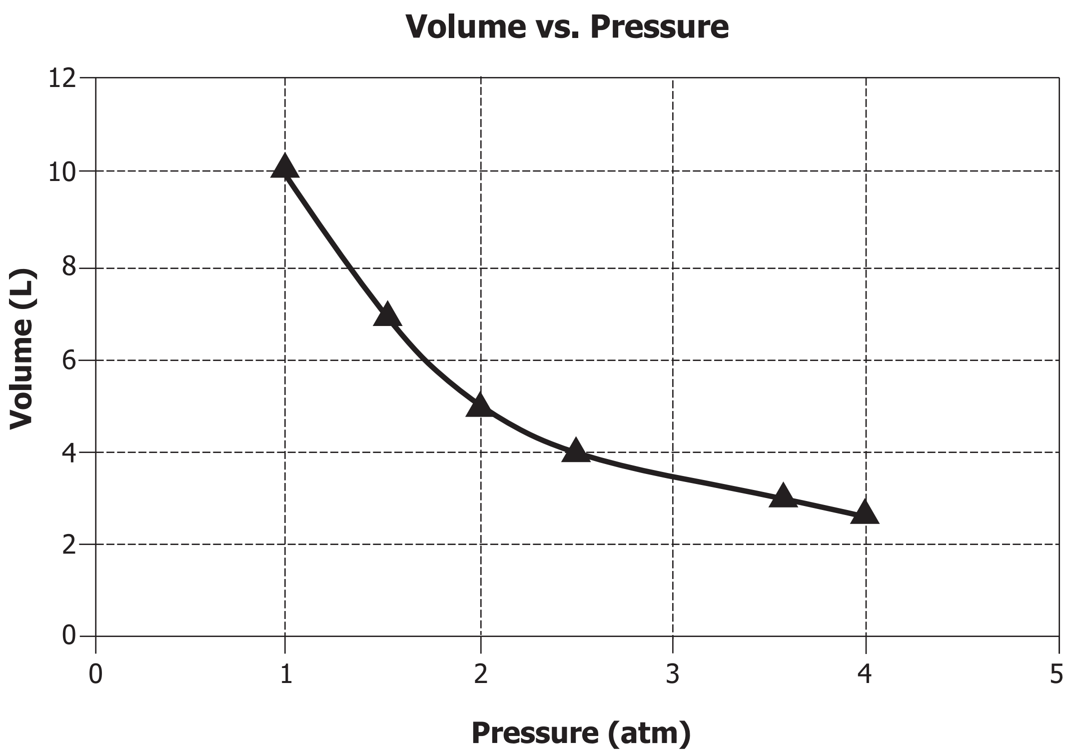
Below you will see six cylinders, each one containing the same type of gas at the same temperature. Each cylinder has a plunger that is compressing the gas to a certain volume, thus creating a certain amount of pressure in each cylinder. After you observe these six cylinders, use the data table to record the volume and pressure of each one. The first set of data has been recorded for you.



Pressure (atm)	Volume (L)
1	10
1.5	7
2	5
2.5	4
3.5	3
4	2.5

Using the data from your table, you can now complete the Pressure-Volume graph below. We will refer to this as Graph 1. The first point has been plotted for you. Once all six points are plotted, draw a smooth curve linking all of the points together.

Graph 1



Use the Pressure-Volume graph (Graph 1) to answer the following questions:

1. What is the relationship between the pressure and the volume of a gas? Give an example from the graph.

   Answer:

   As pressure increases, volume decreases (and vice-versa). For example, at 2.0 atm the volume of gas is 5.0 L, while at 2.5 atm of pressure the volume has decreased to 4.0 L.

2. From the graph, what would the volume of gas be if the pressure were 3.5 atm?

   Answer:

   About 3.0 L.

3. From the graph, what would the volume of gas be if the pressure were 1.5 atm?

   Answer:

   About 7.0 L.

4. From the graph, what would the gas pressure be if the volume were 6.0 L?

   Answer:

   About 1.75 atm.

5. From the graph, what would the gas pressure be if the volume were 3.0 L?

   Answer:

   About 3.5 atm.

6. What would the volume be if the pressure were increased to 5 atm? Hint: You must extrapolate the curve before you can answer the question.

   Answer:

   About 1.5 L.

7. Why is it dangerous to puncture an aerosol can?

   Answer:

   Aerosol cans contain pressurized gases. If you puncture a can, the gas will escape quickly and the can might explode due to the rapid release.

Learning Activity 2.3: The Relationship between Volume and Pressure

1. Use the graph in the lesson to describe the inverse relationship between the volume at 1 atm and 2 atm.

   Answer:

   At 1 atm the volume is 10 L. At 2 atm the volume has decreased to 5 L. This example illustrates an inverse relationship, because as pressure doubles volume is halved.

2. You release a helium-filled balloon into the sky. If the temperature remains constant, what will happen to the volume and pressure of the helium as the balloon rises higher in the sky?

   Answer:

   Since air pressure decreases as elevation increases, the pressure of the helium in the balloon will decrease. At the same time, the volume of helium in the balloon will increase as the space between the gas particles increases.

3. If a gas is compressed from 2 L to 1 L and the temperature remains constant, what will happen to the pressure? Why?

   Answer:

   The pressure will double if the volume is decreased by one-half since there is an inverse relationship between the volume and the pressure of a gas.

4. Use your own words to explain Boyle's Law.

   Answer:

   Boyle's Law states that as the pressure of a sample of gas increases (a fixed number of particles), the volume decreases as long as the temperature is constant.

Learning Activity 2.4: Working with Significant Figures

1. Determine the number of significant digits in each of the following numbers.

5.897	**4**	8.000	**4**	10001	**5**
0.333	**3**	8.001	**4**	0.008000	**4**
7	**1**	0.009	**1**	947.000	**6**
10000	**1**	12000	**2**	10000.0	**6**
10321	**5**	55040	**4**	375000	**3**

2. Complete the following questions respecting the correct number of significant figures in your answers.

   a) 23.1 + 4.77 + 125.39 + 3.581

   Answer:

   To one decimal place, the answer is 156.8 (three significant figures).

   b) 22.101 – 0.9307

   Answer:

   To three decimal places, the answer is 21.170 (five significant figures).

   c) 2.33 x 6.085 x 2.1

   Answer:

   To two significant figures, the answer is 30.

   d) 8432 ¸ 12.5

   Answer:

   To three significant figures, the answer is 674.

Learning Activity 2.5: Using Boyle's Law

1. Use Boyle's Law to calculate the following. Show your work and always use significant figures for your answers.

   a) $V_1 = 2.0$ L, $P_1 = 0.82$ atm, $V_2 = 1.0$ L, $P_2 = ?$

   Answer:

   Only one ratio will result in increased pressure as volume decreases.

   $$\text{new pressure} = 0.82 \text{ atm} \times \left(\frac{2.0 \text{ L}}{1.0 \text{ L}}\right) = 1.6 \text{ atm}$$

   b) $V_1 = 0.55$ L, $P_1 = 740$ mm Hg, $V_2 = 0.80$ L, $P_2 = ?$

   Answer:

   Only one ratio will result in decreased pressure as volume increases.

   $$\text{new pressure} = 740 \text{ mmHg} \times \left(\frac{0.55 \text{ L}}{0.80 \text{ L}}\right) = 510 \text{ mmHg}$$

   c) $V_1 = 4.0$ L, $P_1 = 210$ kPa, $V_2 = 2.5$ L, $P_2 = ?$

   Answer:

   Only one ratio will result in increased pressure as volume decreases.

   $$\text{new pressure} = 210 \text{ kPa} \times \left(\frac{4.0 \text{ L}}{2.5 \text{ L}}\right) = 340 \text{ kPa}$$

2. The volume of a gas at 98 kPa is 250 mL. If the pressure is increased to 180 kPa, what will the new volume be?

   Answer:

   Only one ratio will result in increased pressure as volume decreases.

   $$\text{new volume} = 250 \text{ mL} \times \left(\frac{98 \text{ kPa}}{180 \text{ kPa}}\right) = 140 \text{ mL}$$

3. A gas is placed into a syringe until the pressure is 45.0 kPa. What is the new pressure if

   a) the volume in the syringe is doubled?

      Answer:

      Doubling the volume should decrease the pressure by one-half. The new pressure will be one-half 45.0 kPa or 22.5 kPa.

   b) the volume in the syringe is tripled?

      Answer:

      If the volume in the syringe is tripled, the pressure should decrease to one-third of 45.0 kPa or 15.0 kPa.

   c) the volume is one-third its original volume?

      Answer:

      If the volume is one third its original volume the pressure should increase by three times or 135 kPa.

Learning Activity 2.6: Working with Charles' Law

1. Calculate the final temperature, in degrees Celsius, when the volume of 1.00 L of a gas, at 25.0°C, is doubled to 2.00 L.

   Answer:

   Change temperature to Kelvin.

   $25.0 + 273 = 298\text{ K}$

   Since temperature and volume vary directly, multiply by the ratio that will increase the temperature.

   $$\text{new temperature} = 298\text{ K}\left(\frac{2.00\text{ L}}{1.00\text{ L}}\right) = 596\text{ K}$$

   Change temperature back to Celsius.

   $596\text{ K} - 273 = 323°\text{C}$

2. Calculate the volume of 100.0 mL of a gas if its temperature is doubled from 20.0°C to 40.0°C.

   Answer:

   Change the temperatures to Kelvin.

   20.0°C + 273 = 293 K

   40.0°C + 273 = 313 K

   Because temperature is increasing, multiply by the ratio that will increase the volume.

$$\text{new volume} = 100.0\ \text{mL}\left(\frac{313\ \text{K}}{293\ \text{K}}\right) = 107\ \text{mL}$$

3. You get a balloon at the circus with a volume of 2.50 L at room temperature (25.0°C). You then step outside on a cold winter day (–25.0°C). If the pressure remains constant, what is the balloon's new volume?

   Answer:

   Change the temperatures to Kelvin.

   25.0°C + 273 = 298 K

   −25.0°C + 273 = 248 K

   Since temperature decreases, multiply by ratio that will decrease volume:

$$\text{new volume} = 2.50\ \text{L}\left(\frac{248\ \text{K}}{298\ \text{K}}\right) = 2.08\ \text{L}$$

4. 20.0 mL of a gas at 285 K increases in volume to 32.0 mL. What is the new temperature of the gas?

   Answer:

   The temperature is already in Kelvin.

   Volume increases, so temperature was increased. Multiply by the ratio that will increase temperature.

$$\text{new temperature} = 285\ \text{K}\left(\frac{32.0\ \text{mL}}{20.0\ \text{mL}}\right) = 456\ \text{K}$$

5. Find the new volume if the following changes occur at a constant pressure:
   a) 225 mL of oxygen at 273 K is warmed to 301 K.

      Answer:

      The temperatures are already in Kelvin.

      Temperature increases, so volume will increase. Multiply by the ratio that will increase volume.

$$\text{new volume} = 225 \text{ mL} \left(\frac{301 \text{ K}}{273 \text{ K}} \right) = 248 \text{ mL}$$

   b) 3.0 L of nitrogen is cooled from 90.0°C to −45.0°C.

      Answer:

      Convert temperatures to Kelvin:

      90.0°C + 273 = 363 K

      −45.0°C + 273 = 228 K

      Temperature decreases, so volume will decrease. Multiply by the ratio that will decrease volume.

$$\text{new volume} = 3.0 \text{ L} \left(\frac{228 \text{ K}}{363 \text{ K}} \right) = 1.9 \text{ L}$$

6. Would it be possible for a sample of gas to have a volume of zero? Explain.

   Answer:

   As the volume of a gas decreases, the space between the gas molecules will decrease. Eventually the gas will condense to a liquid, then compress to become a solid. There is a point beyond which the sample cannot be compressed because the molecules actually occupy space.

Learning Activity 2.7: Gay-Lussac's Law

1. A gas in a rigid container has a pressure of 1.00 atm at 25.0°C. What is the pressure of the gas at 50.0°C?

   Answer:

   Convert temperatures to Kelvin:

   25.0°C + 273 = 298 K

   50.0°C + 273 = 323 K

   Temperature increases, so pressure increases. Multiply by the ratio that will increase pressure.

   $$\text{new pressure} = 1.00 \text{ atm}\left(\frac{323 \text{ K}}{298 \text{ K}}\right) = 1.08 \text{ atm}$$

2. A gas in a rigid container has a pressure of 695 torr at 19.0°C. What is the pressure of the gas at −20.0°C?

   Answer:

   Convert temperatures to Kelvin:

   19.0°C + 273 = 292 K

   –20.0°C + 273 = 253 K

   Temperature decreases, so pressure decreases. Multiply by the ratio that will decrease pressure.

   $$\text{new pressure} = 695 \text{ torr}\left(\frac{253 \text{ K}}{292 \text{ K}}\right) = 602 \text{ torr}$$

3. The water vapour in a pressure cooker is at standard pressure (101.3 kPa) at 30.0°C. What is the temperature, in degrees Celsius, when the pressure is 151 kPa?

   Answer:

   Convert temperature to Kelvin:

   30.0°C + 273 = 303 K

   Standard pressure is 101.3 kPa.

   Pressure increases, so temperature increases. Multiply by the ratio that will increase temperature.

   $$\text{new temperature} = 303\text{ K}\left(\frac{151\text{ K}}{101.3\text{ K}}\right) = 452\text{ K}$$

   Convert to Celsius.

   452 – 273 = 179°C

4. The air pressure in a scuba tank is 175 atm at room temperature, 22.0°C. What is the pressure in the tank while diving in water at a temperature of 6.0°C?

   Answer:

   Convert temperatures to Kelvin:

   22.0°C + 273 = 295 K

   6.0°C + 273 = 279 K

   Temperature decreases, so pressure decreases. Multiply by the ratio that will decrease pressure.

   $$\text{new pressure} = 175\text{ atm}\left(\frac{279\text{ K}}{295\text{ K}}\right) = 166\text{ atm}$$

5. You fill your tires on a cool morning (10.0°C) to a pressure of 205 kPa. If you take a long driving trip where the tire temperature reaches 75.0°C, what is the pressure in the tire while you are driving?

   Answer:

   Convert temperatures to Kelvin:

   10.0°C + 273 = 283 K

   75.0°C + 273 = 348 K

   Temperature increases, so pressure increases. Multiply by the ratio that will increase pressure.

   $$\text{new pressure} = 205 \text{ kPa}\left(\frac{348 \text{ K}}{283 \text{ K}}\right) = 252 \text{ kPa}$$

Learning Activity 2.8: The Combined Gas Law

Solve each of the following problems using the ratio method and the Combined Gas Law. Include appropriate units for your final answer.

1. In a laboratory experiment, 85.3 mL of a gas are collected at 24.0°C and 733 mmHg pressure. Find the volume at 760 mmHg pressure and 0.0ºC.

   Answer:

   P_1 = 733 mm Hg

   P_2 = 760 mmHg

   T_1 = 24.0°C + 273 = 297 K

   T_2 = 0.0°C + 273 = 273 K

   V_1 = 85.3 mL

   V_2 = ?

   (**Note:** 760 mmHg is a standard value with 3 significant digits)

   Using ratios, an increase in pressure and a decrease in temperature will decrease the volume.

   $$V_2 = 85.3 \text{ mL} \times \left(\frac{733 \text{ mmHg}}{760 \text{ mmHg}}\right)\left(\frac{273 \text{ K}}{297 \text{ K}}\right) = 75.6 \text{ mL}$$

Use the Combined Gas Law.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(733 \text{ mmHg})(85.3 \text{ mL})}{297 \text{ K}} = \frac{(760 \text{ mmHg})V_2}{273 \text{ K}}$$

Rearrange to solve for V_2.

$$V_2 = \frac{(733 \text{ mmHg})(85.3 \text{ mL})(273 \text{ K})}{(297 \text{ K})(760 \text{ mmHg})}$$

$$V_2 = 75.6 \text{ mL}$$

2. A flexible container has a maximum volume of 15.0 L. If the container is filled to 8.00 L at a pressure of 1.80 atm and a temperature of 17.0°C, at what temperature will the container have a maximum volume at a pressure of 2.90 atm?

Answer:

P_1 = 1.80 atm

P_2 = 2.90 atm

T_1 = 17.0°C + 273 = 290 K

T_2 = ?

V_1 = 8.00 L

V_2 = 15.0 L

Using ratios, increasing pressure and increasing volume will increase the temperature.

$$T_2 = 290 \text{ K} \times \left(\frac{2.90 \text{ atm}}{1.80 \text{ atm}}\right)\left(\frac{15.0 \text{ L}}{8.00 \text{ L}}\right) = 876 \text{ K}$$

Use the Combined Gas Law.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(1.80 \text{ atm})(8.00 \text{ L})}{2.90 \text{ K}} = \frac{(2.90 \text{ atm})(15.0 \text{ L})}{T_2}$$

Rearrange to solve for T_2.

$$T_2 = \frac{(2.90 \text{ atm})(15.0 \text{ L})(290 \text{ K})}{(1.80 \text{ atm})(8.00 \text{ L})}$$

$$T_2 = 876 \text{ K}$$

3. If you have 17.0 L of gas at a temperature of 28.0°C and a pressure of 88.9 atm, what will be the pressure of the gas if you raise the temperature to 68.0°C and decrease the volume to 12.0 L?

   Answer:

   P_1 = 88.9 atm

   P_2 = ?

   T_1 = 28.0°C + 273 = 301 K

   T_2 = 68.0°C + 273 = 341 K

   V_1 = 17.0 L

   V_2 = 12.0 L

   Using ratios, increasing the temperature while decreasing the volume will increase the pressure.

$$P_2 = 88.9 \text{ atm} \times \left(\frac{17.0 \text{ L}}{12.0 \text{ L}}\right)\left(\frac{341 \text{ K}}{301 \text{ K}}\right) = 143 \text{ atm}$$

   Using the Combined Gas Law.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(88.9 \text{ atm})(17.0 \text{ L})}{301 \text{ K}} = \frac{P_2(12.0 \text{ L})}{341 \text{ K}}$$

   Rearrange to solve for P_2.

$$P_2 = \frac{(88.9 \text{ atm})(17.0 \text{ L})(341 \text{ K})}{(12.0 \text{ L})(301 \text{ K})}$$

$$P_2 = 143 \text{ atm}$$

4. A gas that has a volume of 28.0 L, a temperature of 45.0° C, and an unknown pressure has its volume increased to 34.0 L and its temperature decreased to 35.0°C. If the pressure after the change is 2.00 atm, what was the original pressure of the gas?

   Answer:

   $P_1 = ?$

   $P_2 = 2.00$ atm

   $T_1 = 45.0°C + 273 = 318$ K

   $T_2 = 35.0°C + 273 = 308$ K

   $V_1 = 28.0$ L

   $V_2 = 34.0$ L

   When using ratios to find initial values, imagine the final values were the starting values and work backward. The temperature was decreased from 45.0°C to 35.0°C so the starting pressure was greater than the final pressure. The volume was increased, so the initial pressure was greater than the final pressure.

   $$P_1 = 2.00 \text{ atm} \times \left(\frac{34.0 \text{ L}}{28.0 \text{ L}}\right)\left(\frac{318 \text{ K}}{308 \text{ K}}\right) = 2.51 \text{ atm}$$

   Using the Combined Gas Law.

   $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

   $$\frac{P_1(28.0 \text{ L})}{318 \text{ K}} = \frac{(2.00 \text{ atm})(34.0 \text{ L})}{308 \text{ K}}$$

   Rearrange to solve for P_1.

   $$P_1 = \frac{(2.00 \text{ atm})(34.0 \text{ L})(318 \text{ K})}{(28.0 \text{ L})(308 \text{ K})}$$

   $$P_1 = 2.51 \text{ atm}$$

Learning Activity 2.9: The Gas Laws at Work

For each of the following applications, identify the related gas law. Use your course notes as well as other resources, if necessary, to help you.

1. A diver swimming at a depth of 10 m under the water observes that the size of his air bubbles increase in volume as they rise to the water's surface.

   Gas Law: ***Boyle's Law***

2. As a diver swims deeper into the water, added pressure can cause more nitrogen to be absorbed by your blood than normal. If the diver ascends too quickly, the decrease in pressure can cause those dissolved nitrogen bubbles to expand. This can block the flow of blood to critical organs, cause severe joint pain (known as "the bends"), or produce effects similar to alcohol intoxication (called nitrogen narcosis).

   Gas Law: ***Boyle's Law***

3. Nitrogen is pumped into race car tires because it has a more consistent rate of expansion and contraction as the tires heat up on the track and cool down after racing. The expansion and contraction of nitrogen gas creates pressure changes in the tire that are more predictable than when a regular air mixture is used.

   Gas Law: ***Gay-Lussac's Law***

4. "Air blasters" are commercial dusting sprays. They contain a large volume of air that has been compressed into a small space. When the nozzle is pressed, some of the gas escapes from the can.

   Gas Law: ***Boyle's Law***

5. As a plunger is pushed down into a clogged toilet, the volume of air in the plunger becomes less, and the pressure increases. This increase in pressure is transferred through the water in the drain, which breaks apart the clog.

   Gas Law: ***Boyle's Law***

6. As the air in a hot air balloon is heated, its volume expands and the balloon rises, because the heated air is less dense than the cooler air around it.

   Gas Law: ***Charles' Law***

7. On a hot summer day, your friend shakes a cold can of pop before handing it to you. You tap a few times on the lid (to increase the pressure on the can so that more carbon dioxide gas is forced to dissolve into the liquid) and then pop the tab slowly (so that the drop in pressure inside the container is not too sudden).

   Gas Law: ***Boyle's Law***

8. The plunger on a turkey syringe thermometer pops out when the turkey is done, because the volume of air trapped under the plunger increases as the temperature inside the turkey increases.

   Gas Law: ***Charles' Law***

9. An air mattress that is blown up in a cool basement feels firmer when it is moved into a warmer room upstairs.

   Gas Law: ***Charles' Law***

10. Throwing an aerosol can into a fire can cause an explosion.

    Gas Law: ***Gay-Lussac's Law***

11. A typical diver's lungs can hold up to 8 L of air. As a diver descends, the air in the diver's cavities (i.e., lungs, sinuses, and middle ears) decreases in volume while the pressure increases. At a depth of 10 m, the same amount of gas will occupy only half the volume (4 L) that it would at sea level.

    Gas Law: ***Boyle's Law***

12. Scuba shop owners often fill tanks in a water bath where the temperature is approximately the same as wherever the diving will take place. They do this because a full scuba tank will gain about 5 or 6 p.s.i. of pressure for every one degree rise in temperature. A tank of 3000 p.s.i. that is heated to 3500 p.s.i. could easily explode.

    Gas Law: ***Gay-Lussac's Law***

13. Scuba divers carry a small container on their back (0.39 $ft.^3$ or 28.3 L), but have large volumes of breathable air available (80 $ft.^3$ at sea level, enough to fill a phone booth). This is because the air is stored at high pressures, and then released at ambient pressure (1 atm) using a regulator.

    Gas Law: ***Boyle's Law***

14. A football inflated inside and then taken outdoors on a winter day shrinks slightly.

    Gas Law: ***Charles' Law***

15. In a combustion engine cylinder of constant volume, a dramatic increase in temperature will increase the pressure that the gas exerts on the piston. The compression of this piston is what generates the power that moves the car

    Gas Law: ***Gay-Lussac's Law***

16. A slightly overinflated raft explodes after being launched in tropical waters.

    Gas Law: ***Charles' Law***

17. The gauge pressure in a steel-belted automobile tire is higher after travelling over hot asphalt.

    Gas Law: ***Gay-Lussac's Law***

18. When deep sea fish are brought to the surface where the pressure is lower, they die because the volume of gas in their bodies increases, causing bladders, cells, and membranes to pop.

    Gas Law: ***Boyle's Law***

19. In the event of an automobile accident, heat from an ignition source will cause sodium azide, silicon dioxide, and potassium nitrate to produce the 67 L of nitrogen gas that fills an air bag in 0.03 seconds.

    Gas Law: ***Charles' Law***

20. To submerge a submarine, pumps allow water into special tanks and puts air under pressure to decrease its volume. This increases the density of the submarine so that it "dives" down into the water. To surface, the submarine decreases pressure to allow the air to expand back to its original volume and expel the water. The density of the sub then becomes less than that of water and it floats up to the surface.

    Gas Law: ***Boyle's Law***

Grade 11 Chemistry (30S)

Module 3: Chemical Reactions

Module 3: Chemical Reactions

Introduction

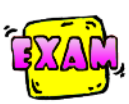
Did you know that there are three different forms of hydrogen? These different variations of atoms of the same element are called isotopes, and you will learn about where they are found in the world around you.

Next, you will learn how to write names and formulas for polyatomic compounds. This will set the stage for the next step—writing and balancing chemical equations. You will see how a balanced chemical equation is actually very similar to a recipe for making your favourite cake!

Moving on through the module, you will discover that a “mole” is not just a small, furry mammal. In fact, the mole is the scientific unit that relates the number of particles in a sample to its mass and is an essential tool in chemistry. Finally, you will use the mole to solve problems involving balanced chemical equations.

You will be assessed on how well you complete the following assignments. They are found within the lessons themselves. Once you have completed the entire module, you will mail your assignments to the Distance Learning Unit. The instructions for doing so are found in the course Introduction.

Writing Your Midterm Examination



You will write the midterm examination when you have completed Module 3 of this course. The midterm examination is based on Modules 1 to 3, and is worth 20 percent of your final mark in the course. To do well on the midterm examination, you should review all the work you complete in Modules 1 to 3, including all the learning activities and assignments. You will write the midterm examination under supervision.

Assignments in Module 3

When you have completed the assignments for Module 3, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 3.1	Working with Isotopes
2	Assignment 3.2	Working with Chemical Compounds
3	Assignment 3.3	Determining Formula Mass
4	Assignment 3.4	Classifying and Balancing Equations
5	Assignment 3.5	Reaction Types
6	Assignment 3.6	Calculating Molar Mass
7	Assignment 3.7	Determining the Volume of a Gas
8	Assignment 3.8	Converting between Mass, Moles, and Number of Particles
9	Assignment 3.9	Determining Empirical and Molecular Formulas from Percent Composition

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

Lesson 1: Atomic Mass and Isotopes (2 hours)

Lesson Focus

SLO C11-3-01: Determine average atomic mass using isotopes and their relative abundance.
Include: atomic mass unit (amu)

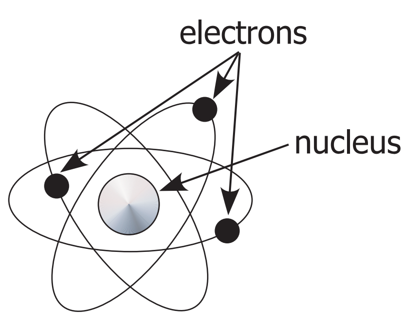
SLO C11-3-02: Research the importance and applications of isotopes.
Examples: nuclear medicine, stable isotopes in climatology, dating techniques...

Lesson Introduction

Red grapes, green grapes, and Concord grapes are all varieties of grapes that you can find at the grocery store. While these are all grapes, their taste, colour, and size are different one from the other. In this lesson, you will learn about the different varieties of atoms, called isotopes. You will also discover the many different ways isotopes are used.

A Review of the Atom

You may recall from earlier science courses that all atoms (except hydrogen) are made of three basic particles: protons, neutrons, and electrons. **Electrons** surround the nucleus, while **protons** and **neutrons** are both found within the nucleus.



You might remember that the **atomic number** of an element is the number of protons in the nucleus of an atom of that element. Each element has a unique number of protons, which is indicated by its atomic or ***Z*** number (the symbol for atomic number is *Z*). Different elements are different because they contain a different number of protons in their nucleus; therefore, we use atomic number to identify an element. Lithium, the element represented by the diagram above, has three protons and therefore also has an atomic number of three. Thus, the number of electrons equals the number of protons in any neutral atom.

In Grade 10 Science, you could have learned that when a given number of negatively charged particles combine with a given number of positively charged particles, an electrically neutral particle is formed.

The number of neutrons in each atom varies, even between atoms of the same element. For example, potassium can exist as three different forms of the same atom. All three atoms contain 19 protons, but one type of potassium atom has 20 neutrons, another 21 neutrons, and yet another has 22 neutrons.

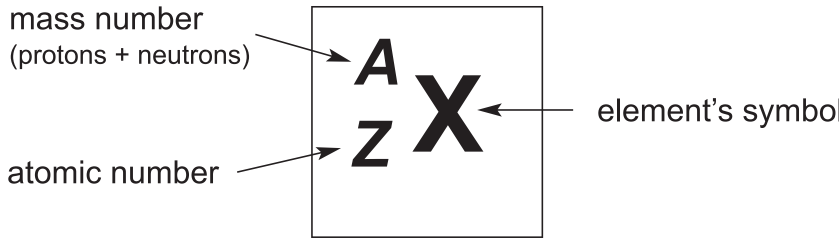
What Is an Isotope?

Atoms that have the same number of protons but differ in their number of neutrons are called **isotopes**. Generally, most elements exist as more than one isotope.

As you would expect, if different isotopes have a different number of neutrons, they will have different masses and different mass numbers. To find the **mass number** (or ***A*** number) of an atom, simply add the number of protons and neutrons found in the nucleus of that atom. Remember that electrons *do* have a mass, but it is so small that you do not consider it for these calculations. If you consider the potassium isotopes mentioned above:

- The isotope containing 19 protons and 20 neutrons will have a mass number of 39 (19 + 20). We call this isotope *potassium-39*.
- The isotope that has 19 protons and 21 neutrons will have a mass number of 40 (19 + 21) and is called *potassium-40*.
- The isotope that has 19 protons and 22 neutrons will have a mass number of 41 (19 + 22) and is called *potassium-41*.

Chemists have designed a symbol for each isotope that includes the element's symbol, its atomic number (*Z*), and its mass number (*A*).



For example, the symbol for potassium-39 would be:

$$^{39}_{19}\text{K} \quad \text{or} \quad ^{39}\text{K}$$

Despite these differences, isotopes are chemically alike, since they share a common number of protons and electrons. It is these two subatomic particles, not neutrons, that are primarily responsible for the chemical behaviour of an element.

Atomic Mass Unit

The masses of individual atoms are expressed as **atomic mass units** (amu, u, or μ). The atomic mass unit is defined as $\frac{1}{12}$ the mass of a carbon-12 atom. This means a proton or a neutron has mass equal to *approximately* one atomic mass unit. Carbon-12 was an arbitrary choice, meaning there was no special reason for selecting it as the point of comparison. It was likely used because carbon is a very common element and the value for the atomic masses of many other atoms is then very nearly a whole number. Through experimentation, the atomic mass unit of carbon was determined to be about 1.66×10^{-24} g, which is an extremely small unit of mass.

Did you know? There is an atomic mass unit called the **Dalton (Da)**, named after John Dalton. Dalton was the first scientist to suggest using the mass of one atom of hydrogen as the atomic mass unit. The Dalton is not technically an SI unit, but it is used in biochemistry and molecular biology to measure large protein molecules.

1 u = 1 Da

1 Da = $1.66053886 \times 10^{-24}$ g

In nature, many elements can be found as a mixture of two or more isotopes. You have probably noticed that the atomic mass shown for each element on the periodic table is rarely a whole number. This is because it is actually an average mass of all of the isotopes of that element. Every isotope of an element has a definite mass and a **relative natural abundance**. The relative abundance of an isotope is the fraction of each isotope found in an average sample of the element. For example, all three isotopes of potassium can be found in bananas according the following abundances:

Potassium-39 (20 neutrons)	93.1%
Potassium-41 (22 neutrons)	6.88%
Potassium-40 (21 neutrons)	0.02%

Calculating Average Atomic Mass Using Relative Abundance

Example 1

Magnesium exists as three isotopes: magnesium-24, magnesium-25, and magnesium-26. In an average sample of magnesium, the relative abundance breakdown is as follows:

Magnesium-24	78.99%
Magnesium-25	10.00%
Magnesium-26	11.01%

If the atomic mass of magnesium-24 is 23.985 amu, magnesium-25 is 24.986 amu, and magnesium-26 is 25.982 amu, calculate the average atomic mass of magnesium.

Solution

Perhaps you may have noticed that the actual atomic mass of magnesium-24 is not 24 amu, but 23.985 amu. This is because a single amu is not exactly equal to the mass of a proton or a neutron.

The **average atomic mass** is the *weighted average* of the relative abundances of each isotope. For magnesium, that would be:

$$\left(\frac{78.99\%}{100\%}\right)23.985\ \mu + \left(\frac{10.00\%}{100\%}\right)24.986\ \mu + \left(\frac{11.01\%}{100\%}\right)25.982\ \mu = 24.305\ \mu$$

If you complete each set of calculations separately, you should have come up with the answer of 24.306 based on your rounding. If you look at the periodic table, you will see that the average atomic mass of magnesium is about 24.3 amu. You will also notice that the average mass of the magnesium is closer to 24 amu than to 25 amu or 26 amu. This is because magnesium-24 has a greater abundance than the other two isotopes.

We can relate this to students' marks. When you think of a weighted average, you might see a similarity in the way your marks are determined in this, or any, course. Tests may be worth 30%, quizzes may be worth 15%, labs and assignments may be 25%, and the final examination 30%. Tests and the final examination would have a greater effect on an individual student's mark because of their higher weights.

How Are Isotopes Used?

Isotope Applications

1. **Radioactive Tracers in Medical Diagnosis**
   - ^{131}I can be used to image the thyroid, heart, lungs, and liver, and to measure iodine levels in blood.
   - ^{24}Na (a beta emitter with a half-life of 14.8 h) injected into the bloodstream as a salt solution can be monitored to trace the flow of blood and detect possible constrictions or obstructions in the circulatory system.
   - PET (positron emission tomography) scans use ^{15}O in $H_2{}^{15}O$ and ^{18}F bonded to glucose to measure energy metabolism in the brain.
2. **Radioactive Isotopes in Medical Treatment**
   - Implants of ^{198}Au or mixtures of ^{90}Sr and ^{90}Y have been used to destroy pituitary and breast cancer tumours.
   - Gamma rays from ^{60}Co are used to destroy brain tumours.
3. **Oxygen Isotopes in Climatology and Geology**

   Stable isotopes such as ^{16}O and ^{18}O are used to identify global temperatures in the distant past. This can be done by determining the ratio of ^{18}O to ^{16}O in ice cores extracted from Earth's polar caps or from sediment cores exhumed from the ocean floor. There is a correlation between an excess of the light isotope, ^{16}O, in precipitation and global temperatures.

   When ice sheets grow in the polar regions during glacial periods, they incorporate water that has been evaporated in the low latitudes and carried to the poles in the form of water vapour, which is then precipitated as snow.

Evaporation favours the light isotope of oxygen, ^{16}O, for reasons of simple kinetics, and so polar ice has proportionally more ^{16}O than the seawater left behind when evaporation rates are high (warmer periods). This means that newly deposited ocean sediments will have a greater abundance of the *heavier isotope*, ^{18}O, when world temperatures are higher than average. Therefore:

- When world ice volume increases during a glacial stage, the heavier isotope, ^{18}O, *decreases* in polar ice and snow.
- When ice volume shrinks during warming (interglacial) periods, such as we have right now, the abundance of ^{18}O *increases* in the world's oceans. This shows up in decreased ^{18}O content in polar ice.

This *isotopic signature* can be preserved in certain shelled animals such as marine Foraminifera. These tiny bottom-dwellers secrete a silicate shell that carries the ratio of $^{18}O/^{16}O$ consistent with what that ratio was in the seawater around it. "Heavy shell" = more ice on the planet. This makes for a very effective *paleothermometer* that can be used to correlate ocean temperatures, world climate, and sea-ice volumes.

If you would like to learn more about how isotopes of oxygen are used in climate studies, see <http://earthobservatory.nasa.gov/Features/Paleoclimatology_OxygenBalance/>.

4. Carbon and Hydrogen Isotopes in Atmospheric Nuclear Tests

Nuclear weapons tests put large and detectable amounts of certain radioactive isotopes into the atmosphere. After the near-elimination of nuclear bomb testing due to the Limited Test Ban Treaty in 1963, the carbon-14 (^{14}C) concentration in the atmosphere began decreasing immediately. Anybody born *before* 1965 or so possesses a significantly higher concentration of ^{14}C than someone born after atmospheric nuclear testing ended. Thus, we can tell how old many living organisms are (including humans) based on the recent history of ^{14}C content in the atmosphere. Such sources of radiogenic isotopes are often described as coming from *anthropogenic* (human-generated) activities.

A detailed discussion of tritium in the atmosphere can be found at the following websites:

U.S. Environmental Protection Agency: Radiation Information:
www.epa.gov/radiation/radionuclides/tritium.html

U.S. Geological Survey:
wwwrcamnl.wr.usgs.gov/isoig/period/h_iig.html

The U.S. Geological Survey also provides information related to periods of atmospheric nuclear testing.

5. **Isotopes in Dating Techniques**

- Carbon-14, with a half-life of 5730 years, is used to determine the age of bones discovered at archeological sites because the ^{14}C continues to decay over the years, whereas the amount present in the atmosphere is constant. The maximum age of an object for dating purposes using ^{14}C is about 24 000 years, whereas a long-lived isotope such as ^{238}U can be used to date materials up to 4.5×10^9 years old.
- Uranium-238 and lead-206 are commonly used to date very ancient objects such as minerals contained within rock samples.

Learning Activity 3.1

What Is an Isotope?

1. Identify the numbers of protons, neutrons, and electrons in a neutral atom of each of the following:

   a) $^{235}_{92}U$

   b) $^{226}_{88}Ra$

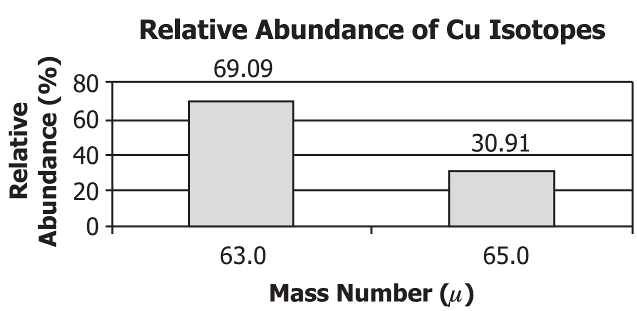
continued

Learning Activity 3.1 (continued)

2. Complete the following table to calculate the average atomic mass of each element.

Element	Symbol	Mass Number	Mass (μ)	Relative Abundance (%)	Average Atomic Mass (μ)
Carbon	C–12	12	12.000	98.98	
	C–13	13	13.003	1.11	
Silicon	Si–28	28	27.977	92.21	
	Si–29	29	28.976	4.70	
	Si–30	30	29.974	3.09	

3. Define the term *isotope*. Explain how an element's atomic mass is related to the abundances of its different isotopes.
4. Using the graph below, calculate the average atomic mass of copper.



5. There are three isotopes of silicon, having the mass numbers of 28, 29, and 30. The atomic mass of silicon is 28.068 amu. Which of the isotopes is the most abundant? Which would be the least abundant?

Lesson Summary

In this lesson, you learned that isotopes are atoms of the same element that differ in mass because they have different numbers of neutrons. You also learned how to calculate the relative (or percent) abundance of an isotope in an average sample of a particular element. Finally, you discovered some of the many uses of isotopes in everyday life. In the next lesson, you will start writing names and formulas for polyatomic compounds.

Notes

Assignment 3.1

Working with Isotopes *(12 marks)*

1. Boron has two isotopes—boron-10 and boron-11. Which is more abundant, given that the atomic mass of boron is 10.81? Explain your reasoning. *(2 marks)*

   ____

   ____

2. Three isotopes of oxygen are oxygen-16, oxygen-17, and oxygen-18. Write the symbol for each, including the atomic number and mass number. *(3 marks)*

3. Complete the following table to calculate the average atomic mass of each element. Show your calculations. *(7 marks)*

Element	**Symbol**	**Mass Number**	**Mass (μ)**	**Relative Abundance (%)**	**Average Atomic Mass (μ)**
Helium	He–3	3	3.0160	0.00010	
	He–4	4	4.0026	99.9999	
Oxygen	O–16	16	15.995	99.759	
	O–17	17	16.995	0.0370	
	O–18	18	17.999	0.204	

NOTES

Lesson 2: Polyatomic Compounds (2 hours)

Lesson Focus

SLO C11-3-03: Write formulas and names for polyatomic compounds using International Union of Pure and Applied Chemistry (IUPAC) nomenclature.

Lesson Introduction

Chemistry has its own language. Chemists all over the world communicate in this language to describe the millions of known compounds. This communication depends on a standard system of naming and writing of formulas for compounds, called the IUPAC system (the acronym stands the International Union of Pure and Applied Chemistry). In this lesson, you will build on your knowledge of naming compounds, a skill that will later help you write chemical formulas and fully understand chemical reactions.

Chemical Formulas

In Grade 10 Science, you may have learned some important foundations of chemistry that will help you write the formulas and names for polyatomic compounds. Therefore, this lesson begins with a review of some of these concepts.

A **chemical formula** is a shorthand method to represent compounds. It uses the elements' symbols and subscripts, and gives the following information:

- The different elements in the compound represented by their symbols.
- The number of atoms of each element in the compound represented by subscripts.

Study the two examples shown below.

Figure 1

element symbols

$$H_2O$$

subscript

Figure 2

element symbols

$$Ca_3(PO_4)_2$$

subscripts

In Figure 1 above, the compound water (H_2O) contains hydrogen (H) and oxygen (O) atoms. The subscripts that follow each element indicate the number of atoms of that element in the compound. The subscript following the symbol for hydrogen is 2, indicating that there are two hydrogen atoms in each water molecule. Notice that there is no subscript following the symbol for oxygen; this shows that there is only one atom of oxygen in a water molecule. Chemists do not write the number one as a subscript when only one atom of that kind of element exists in the compound.

Figure 2 above shows a more complex chemical formula. The compound $Ca_3(PO_4)_2$ contains calcium (Ca), phosphorous (P), and oxygen (O). The subscript after the symbol for calcium indicates that there are three atoms of calcium in one formula unit. The subscript "2" outside of the bracket indicates that every subscript inside the bracket is to be multiplied by two. This means that there are $2 \times 1 = 2$ atoms of phosphorous and $2 \times 4 = 8$ atoms of oxygen in each formula unit of $Ca_3(PO_4)_2$.

Ionic Compounds

You might recall from Grade 10 Science that an ion is a charged particle that forms when a neutral atom gains or loses electrons. Positively charged ions, called **cations**, are formed when an atom loses one or more electrons. A negative ion, called an **anion**, is formed when an atom gains one or more electrons.

Ionic compounds are formed when two or more oppositely charged ions (usually a metal and a non-metal) are attracted to each other. This chemical attraction is called a **chemical bond**. An **ionic bond** is formed when a negatively charged ion is attracted to a positively charged ion. Ions combine together so that their charges add to zero. You will find examples of ions in *Appendix C: Names, Formulas, and Charges of Some Common Ions* and *Appendix D: Common Ions*, found at the end of the module.

Here are some examples of ionic compounds:

$NaCl$ — sodium chloride

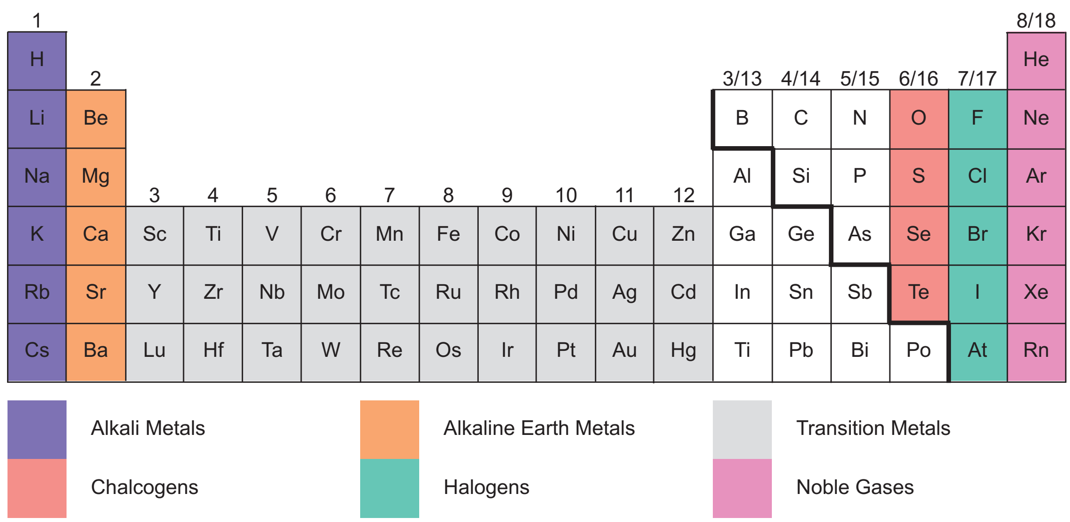
Fe_2O_3 — iron (III) oxide

$CuSO_4$ — copper (II) sulfate

$Ca_3(PO_4)_2$ — calcium phosphate

Predicting the Charge of an Ion

The periodic table groups atoms according to their properties. The periodic table below shows the names of several groups, or families, that you will learn more about throughout this course. Remember that metals are found to the left of the dividing line (the "staircase"), while non-metals are found to the right of the line.



If you do not know the charge of an ion, you can use the periodic table to predict the charge. For example, the alkali metals are found in the first group or family. This means they will each tend to lose 1 electron to produce ions with a 1^+ charge. The outermost shell of electrons is now empty, but the next outermost shell is full. The alkali metal now resembles the closest noble gas and is considered to be stable. Remember that an outermost shell of electrons is stable when it contains eight electrons (or two electrons, in the case of the first shell only). The alkaline earth metals, group 2, tend to form ions with a 2^+ charge. Metals on the left-hand side of the periodic table tend to lose their valence electrons, leaving a complete octet of electrons in their next-lowest energy level.

Interestingly, many of the transition metals, because of their electron arrangement, tend to form more than one ion charge. The reason for this is beyond the scope of this course. You can do an Internet search to find the answer if you are interested in learning more about multiple oxidation states.

On the non-metal side of the table, the chalcogens occupy group 6 (16). This indicates that they have six electrons in their outermost shell and will need to gain two to form a stable group of eight electrons. Therefore, the chalcogens tend to form 2^- ions, while the halogens, group 7 (17), tend to form 1^- ions. What about the noble gases (group 8/18)? Since they are already stable, they do not need to gain or lose electrons, and therefore do not carry a charge.

Naming Binary Ionic Compounds

A binary compound contains two different kinds of elements, although there can be more than one atom of each of those elements. **Binary ionic compounds** usually contain one positively charged metal ion combined with one negatively charged non-metal ion.

When naming an ionic compound from its formula, follow the rules below.

Example 1

Write the name of NaCl.

Step 1: *Name the first element.*

Na = sodium

Step 2: *Name the second element and change the ending to "–ide."*

Cl = chlorine = chloride

The name of the compound is sodium chloride.

Example 2

Write the name of Mg_3P_2.

Step 1: *Name the first element.*

Mg = magnesium

Step 2: *Name the second element and add "–ide."*

P = phosphorous = phosphide

The name of the compound is magnesium phosphide.

Writing Binary Ionic Formulas

You will need the ion charts found in Appendix C and Appendix D to write the formulas of ionic compounds. Remember that formulas contain the *symbols* of the elements involved as well as *subscripts* indicating the number of atoms of each element. Here are the key points to remember when writing the formulas for binary ionic compounds:

- The formula must have the *cation* listed *first*, followed by the anion.
- The *sum* of the charges of the ions must be *zero*. That is, the number of positive charges must equal the number of negative charges.
- You *cannot* change the charge of the ions to make the ion charges equal zero. *You can only modify the subscripts.*

Steps for Writing Formulas Using the Lowest Common Multiple

Step 1: *Write the symbols and charges for the ions involved.*

Step 2: *Determine the lowest common multiple between the charges, and add whatever subscripts are needed to balance the charges.* That is, the number of positive charges must equal the number of negative charges.

Example 3

Write the formula for aluminum oxide.

Step 1: *Write the symbols and charges for the ions involved.*

Aluminum is Al^{3+} and oxide is O^{2-}.

Step 2: *Determine the lowest common multiple between the charges, and add whatever subscripts are needed to balance the charges.*

How can you balance a 3^+ charge and a 2^- charge? You must find the lowest common multiple of the charges. In this case, the lowest common multiple of 3 and 2 is 6. To get 6 positive charges, you would need 2 aluminum ions ($2 \times 3^+ = 6^+$), and to get 6 negative charges, you would need 3 oxygen ions ($3 \times 2^- = 6^-$). The last step is to make these multipliers the subscripts for each element.

Therefore, the formula for aluminum oxide is Al_2O_3.

Metal Ions with Multiple Charges

Most of the transition metals have more than one possible ion charge. Here are some examples:

Ion	Possible Ion Charges
Copper	1^+, 2^+
Iron	2^+, 3^+
Cobalt	2^+, 3^+
Chromium	2^+, 3^+
Lead	2^+, 4^+
Tin	2^+, 4^+

In 1919, a German chemist named Alfred Stock (1876–1946) suggested using Roman numerals to indicate the charge of the ions. Prior to this, the ions were given different names based on their charge. For example, the Cu^+ ion was called "cuprous" and the Cu^{2+} ion was called "cupric." The Fe^{2+} ion was "ferrous" and the Fe^{3+} ion was "ferric," Since the charges were not always the same, when to use the "–ic" and "–ous" endings caused some confusion, especially for chemistry students. Today, the **Stock naming system** uses Roman numerals following the metal ion's name to indicate an ion's charge. For example,

Cu^+ = Copper (I)

Cu^{2+} = Copper (II)

Fe^{2+} = Iron (II)

Fe^{3+} = Iron (III)

Example 4

Write the formula for iron (III) chloride.

Step 1: *Write the symbols and charges for the ions involved* in the usual order, the cation first followed by the anion.

iron (III) is Fe^{3+} and chloride is Cl^-.

Step 2: *Determine the lowest common multiple between the charges, and add whatever subscripts are needed to balance the charges.*

How can you balance a 3^+ charge and a 1^- charge? You must find the lowest common multiple of the charges. In this case, the lowest common multiple of 3 and 1 is 3. To cancel the 3^+ charge associated with one iron ion, you would need three chloride ions. The last step is to make these numbers the subscripts for each.

The formula for iron (III) chloride is $FeCl_3$.

Example 5

Write the formula for lead (IV) sulfide.

Step 1: *Write the symbols and charges for the ions involved* in the usual order, the cation first followed by the anion.

lead (IV) is Pb^{4+} and sulfide is S^{2-}.

Step 2: *Determine the lowest common multiple between the charges, and add whatever subscripts are needed to balance the charges.*

How can you balance a 4^+ charge and a 2^- charge? You must find the lowest common multiple of the charges. In this case, the lowest common multiple for 4 and 2 is 4. To cancel the 4+ charge associated with one lead ion, you would need two sulfur ions ($2 \times 2^- = 4^-$). The last step is to make these numbers the subscripts for each.

Therefore, the formula for lead (IV) sulfide is PbS_2.

Naming Compounds Having Metal Ions with Multiple Charges

You will be naming compounds with multiple charges using the Stock system. The same rules (which you already know) will apply, but you must also determine the charge on the metal ion to correctly name the compound. How can you determine the charge on the metal ion? Let's look at balancing the charges for $CoBr_2$ as an example.

The charge for Br is 1^-. If there are two bromine ions, then the total negative charge would be 2^-. This negative charge would be cancelled by a 2^+ charge on the one cobalt ion. Since the overall net charge of $CoBr_2$ is then zero, you can confirm that the charge of cobalt is in fact 2^+.

In the last step, you can write the name of the compound, indicating the charge of cobalt (2^+) using Roman numerals:

cobalt (II) bromide

You can also try using a table to help you determine the charges of each ion:

Co^{2+}	Br^- Br^-
2^+	2^-

Example 6

Write the name for Fe_2O_3.

Step 1: *Determine the original charges of the individual ions.*

You already know the charge for O is 2^-. If there are three oxygen atoms, then the total charge would be 6^- ($3 \times 2^- = 6^-$). This negative charge must be cancelled by iron. Since there are two atoms of iron and they must have a total charge of 6^+, the charge of each individual ion must be 3^+ ($6^+ \div 2 = 3^+$).

Therefore, $Fe^{3+}O^{2-}$.

Step 2: *Write the name of the compound using Roman numerals following the metal ion's name to indicate its charge.* (Don't forget that you can find the names of ions using Appendix C and Appendix D at the end of the course.)

iron (III) oxide

Learning Activity 3.2

Naming and Writing Formulas for Ionic Compounds

1. Name the following binary ionic compounds:
   a) NiI_2
   b) MgS
   c) K_3N
   d) $FeBr_2$
   e) $CaCl_2$
2. Write formulas for the following binary ionic compounds:
   a) aluminum oxide
   b) tin (IV) sulfide
   c) barium nitride
   d) chromium (III) chloride
   e) magnesium oxide

Writing Names for Compounds Containing Polyatomic Ions

Some ions are composed of several atoms joined covalently. These are called polyatomic ions (*poly* = many). Examples of common polyatomic ions are in the table below. There are many others in Appendix C and Appendix D, found at the end of the course.

Some Common Polyatomic Ions and Their Names		
Ion	**Name**	**Charge**
NH_4^+	ammonium	1+
SO_4^{2-}	sulfate	2–
PO_4^{3-}	phosphate	3–
NO_3^-	nitrate	1–
OH^-	hydroxide	1–
CO_3^{2-}	carbonate	2–
$C_2H_3O_2^-$	acetate	1–

Although polyatomic ions have more than one atom, we will name polyatomic ionic compounds like binary compounds. In other words, you will treat polyatomic ions as though they were one single ion. This means that the charge for polyatomic ions is for the whole group of atoms, not just for the atom that is written last. *Do not* change the subscripts of polyatomic ions; if you do, you change the identity of these ions.

When indicating the presence of more than one polyatomic ion in a compound, we use parentheses around the polyatomic ion, followed by its subscript. For example, the compound $Al(C_2H_3O_2)_3$ has one aluminum ion and three acetate ions. Placing the acetate ion in parentheses and following it with the subscript 3 identifies that 3 acetate ions are required to balance the charge of one aluminum ion. This also helps you avoid confusion as to which atoms belong to which polyatomic ion.

Example 7

Write the name for KNO_3.

Step 1: *Identify the cation and the anion.*

How do we figure out what group of atoms represents the cation, and what group of atoms represents the anion? Except for the ammonium (NH_4^+) ion, all cations in this course will be monoatomic (meaning the positive ion will consist of only one type of element). K^+, then, will be the cation. It only has one possible charge, so we don't need to use a Roman numeral in the formula. The name of the K^+ ion is potassium. NO_3^- will be the anion. It is called the nitrate ion.

Step 2: *Write the name of the cation first, followed by the anion.*

The name of the compound is potassium nitrate.

Example 8

Write the name of $Cu_3(PO_4)_2$.

Step 1: *Identify the anion and the cation.*

Cu will be the cation, and PO_4 will be the anion. Because copper is an ion with more than one possible charge, we must look to the anion to determine its value.

The previous chart tells us that the anion is the phosphate ion (PO_4^{3-}). The parentheses in the formula followed by the number 2 indicate there are 2 phosphate ions in this compound. If phosphate has a charge of $-$, the total charge will be $2 \times 3^- = 6^-$.

The total charge of the anions is 6^-, so the charges of all the cations must add up to 6^+ for there to be a net charge of zero. There are three copper ions in the formula, and $6^+ \div 3 = 2^+$, so the charge on the copper ion is 2^+.

Step 2: *Write the name of the cation first, followed by name of the anion.*

Copper is one of the ions with more than one possible charge, so you must use a Roman numeral to indicate the 2^+ charge. The name of the compound is then copper (II) phosphate.

Example 9

Write the name of NH_4SCN.

Step 1: *Identify the cation and the anion.*

The cation is NH_4^+, or ammonium, while the anion is SCN^-, thiocyanate.

Step 2: *Write the name of the cation first, followed by the name of the anion.*

The name of the compound is ammonium thiocyanate.

Writing Formulas for Compounds Containing Polyatomic Ions

Writing formulas for compounds containing polyatomic ions is similar to writing formulas for binary compounds.

Step 1: *Write out the symbols and charges for the ions involved.*

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.* Be sure to include brackets around polyatomic ions when they have a subscript.

Example 10

Write the formula for sodium perchlorate.

Step 1: *Write out the symbols and charges for the ions involved.*

Sodium is Na^+ and perchlorate is ClO_4^-.

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.*

The sodium is a 1^+ and the perchlorate is 1^-. The lowest common multiple here is 1, so the charges balance with 1 ion each.

The formula for sodium perchlorate is then $Na(ClO_4)$ or $NaClO_4$.

Example 11

Write the formula for iron (III) cyanide.

Step 1: *Write out the symbols and charges for the ions involved.*

Iron (III) is Fe^{3+} and cyanide is CN^{-}.

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.* You may find the following chart method to be helpful:

Fe^{3+}	CN^{-} CN^{-} CN^{-}
3^{+}	3^{-}

Three CN^{-} atoms were needed to balance the charge of the iron atom. The formula for iron (III) cyanide is then $Fe(CN)_3$.

Be sure to put parentheses around the cyanide ion to indicate there are 3 CN^{-} ions. If the formula was written as $FeCN_3$ instead, the formula would indicate that there was one Fe, one C, and 3 Ns, *not* one Fe and 3 CN^{-} ions as there should be.

Example 12

Write the formula for lead (II) sulfate.

Step 1: *Write out the symbols and charges for the ions involved.*

Lead (II) is Pb^{2+} and sulfate is SO_4^{2-}.

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.*

Pb^{2+}	SO_4^{2-}
2^{+}	2^{-}

Only one of each ion is required to balance out the charges, so the formula for lead (II) sulfate is $PbSO_4$.

Example 13

Write the formula for potassium hydrogen phosphate.

Step 1: *Write out the symbols and charges for the ions involved.*

Potassium is K^+ and hydrogen phosphate is HPO_4^{2-}.

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.*

K^+ K^+	HPO_4^{2-}
2^+	2^-

The formula for lead (II) sulfate is $K_2(HPO_4)$ or K_2HPO_4.

Learning Activity 3.3

Writing Formulas for Polyatomic Ions

1. Write the correct formulas for these polyatomic compounds:
   a) Sodium hypochlorite
   b) Strontium nitrate
   c) Cobalt (III) sulfite
   d) Copper (II) phosphate
   e) Zinc hydroxide
2. Name the following polyatomic ionic compounds:
   a) $K_2Cr_2O_7$
   b) $Ni(OH)_2$
   c) $AgNO_3$
   d) Na_3PO_4
   e) $Al_2(SiO_3)_3$

Naming Binary Covalent Compounds

Non-metals tend to combine chemically by sharing electron pairs. These bonds are known as **covalent bonds**. Neutral compounds made of atoms joined covalently are called **molecular** or **covalent compounds**.

We name covalent compounds differently than ionic compounds. We must indicate the number of each element by adding a Greek prefix in front of the element's name.

The prefixes are:

one = mono

two = di

three = tri

four = tetra

five = penta

six = hexa

seven = hepta

eight = octa

nine = nona

ten = deca

Naming Covalent Compounds

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element.* You can omit "mono–" if there is only one of that element in the compound. For example, NO is nitrogen monoxide, but N_2O is dinitrogen monoxide.

Step 2: *Name the second element, but use an "–ide" ending.* Use prefixes to indicate the number of that element (including mono).

Step 3: *Write the name of the compound.*

There are two exceptions to the naming rules. The common names for the following compounds are used instead of their IUPAC names:

H_2O = water

NH_3 = ammonia

Example 14

Write the name for CO_2.

Note: This is a covalent compound since it is made of two non-metal atoms.

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element.*

There is only one carbon atom. You can omit the "mono–" for the first element, so the first part of the name is carbon.

Step 2: *Name the second element. Use prefixes and end the name with "–ide."*

The second element is oxygen. There are two oxygen atoms, so the second part of the name is dioxide.

Step 3: *Write the name of the compound.*

The name of CO_2 is carbon dioxide.

Example 15

Write the name for N_2O_4.

This is another covalent compound, as both nitrogen and oxygen are non-metals.

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element.*

There are two nitrogen atoms, so the first part of the name is dinitrogen.

Step 2: *Name the second element. Use prefixes and end the name with "–ide."*

The second element is oxygen. There are four oxygen atoms, so the second part of the name is tetraoxide.

Step 3: *Write the name of the compound.*

The name of N_2O_4 is then dinitrogen tetraoxide.

Writing Formulas for Binary Covalent Compounds

Writing formulas for binary covalent compounds involves the following steps.

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.* Do not reduce the subscripts for covalent compounds!

Step 3: *Write the name of the compound.*

Example 16

Write the formula for dinitrogen monoxide.

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

The symbol for nitrogen is N and the prefix "di" stands for a subscript of 2.

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.*

The symbol for oxide (oxygen) is O and the prefix "mono" stands for a subscript of 1.

Step 3: *Write the name of the compound.*

The formula for dinitrogen monoxide is then N_2O.

Example 17

Write the formula for sulfur hexafluoride.

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

The symbol for sulfur is S and the lack of a prefix identifies that its subscript is 1.

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.*

The symbol for fluoride is F and the prefix "hexa" stands for a subscript of 6.

Step 3: *Write the name of the compound.*

The formula for sulfur hexafluoride is then SF_6.

Diatomic Molecules

Some elements do not exist as single atoms. These elements exist as pairs of atoms joined covalently, and are called **diatomic molecules**. The elements that exist as diatomic molecules are hydrogen gas (H_2), oxygen gas (O_2), nitrogen gas (N_2), fluorine gas (F_2), chlorine gas (Cl_2), bromine gas or liquid bromine (Br_2), solid iodine (I_2) and solid astatine (At_2). When oxygen gas, hydrogen gas, etc. is used, the formula will be O_2, H_2, etc.

Some students find it helpful to remember the word *HOFBrINCl*, since the symbols of all the diatomic elements (except astatine) are represented in the acronym.

$\boxed{\text{H}}$ ydrogen
$\boxed{\text{O}}$ xygen
$\boxed{\text{F}}$ luorine
$\boxed{\text{B}}$ romine
$\boxed{\text{I}}$ odine
$\boxed{\text{N}}$ itrogen
$\boxed{\text{C}}$ hlorine

You can also use the "7+1" method to help you remember the diatomic elements. Using your periodic table (Appendix E at the end of the course), you can form the number 7, starting with nitrogen (N) and working towards fluorine (F). Now trace the backbone of the number seven going all the way down to iodine (I). All the elements that make up the number seven are diatomic elements. Since hydrogen is also a diatomic element, we say "+1" to help you remember to consider it as well. Recall that 1 is the atomic number of hydrogen.

Learning Activity 3.4

Naming and Writing Formulas for Covalent Compounds

1. Name the following covalent compounds:
   a) NI_3
   b) CO
   c) SF_6
   d) O_2
   e) I_2
2. Write the formulas for the following covalent compounds:
   a) Silicon tetrafluoride
   b) Diphosphorous pentaoxide
   c) Tetraarsenic decaoxide
   d) Hydrogen
   e) Chlorine

Lesson Summary

In this lesson, you reviewed how to name binary ionic compounds whereby the cation is named first, followed by the anion with the "-ide" ending. You also reviewed writing the formula for an ionic compound using subscripts to balance the charges of the ions.

Next, you were introduced to the Stock system. For metal ions with more than one possible ion charge, the Stock naming system uses Roman numerals after the metal ion's name to indicate the charge of the ion.

Finally, you expanded upon your knowledge of polyatomic ions, a group of atoms that share a single charge. You learned how to name and write formulas for binary compounds containing polyatomic ions.

In the next lesson, you will do further work with the atomic mass unit and use this to calculate the mass of compounds.

NOTES

Assignment 3.2

Working with Chemical Compounds *(10 marks)*

1. Name the following binary ionic compounds. *(1 mark)*

   a) Pb_3N_4

   ____

   b) CuCl

   ____

2. Write the formulas for the following covalent compounds. *(1 mark)*

   a) iron (III) phosphide

   ____

   b) boron bromide

   ____

3. Name the following compounds. *(2½ marks)*

   a) Ag_2CrO_4

   ____

   b) $PbCO_3$

   ____

   c) $Sn(SO_4)_2$

   ____

continued

d) $Ba(NO_2)_2$

e) $Ca(C_2H_3O_2)_2$

4. Write the correct formulas for the following polyatomic ionic compounds. *(2½ marks)*

a) Potassium carbonate

b) Ammonium dichromate

c) Iron (III) phosphate

d) Lithium hydrogen sulfide

e) Calcium perchlorate

continued

Assignment 3.2: Working with Chemical Compounds (continued)

5. Name the following covalent compounds. *(1½ marks)*
   a) P_4O_{10}

   ____

   b) N_2O_5

   ____

   c) F_2

   ____

6. Write the formulas for the following covalent compounds. *(1½ marks)*
   a) Dichlorine heptaoxide

   ____

   b) Carbon disulfide

   ____

   c) Bromine

   ____

NOTES

Lesson 3: Atomic Mass Units (1 hour)

Lesson Focus

SLO C11-3-04: Calculate the mass of compounds in atomic mass units.

Lesson Introduction

In Lesson 1, we used the analogy of the grocery store and different varieties of grapes to introduce isotopes. Regardless of what type of grape you might like, the measurement scale used in the produce section is the kilogram (kg). This is because it is easier to measure in less precise units (for example, by 0.5 kg as opposed to 0.446 kg). Likewise in chemistry, there are preferred units of measurement for the atom, the molecule, and any chemical formula.

In this lesson, you will calculate the mass of compounds in atomic mass units. To help you understand this concept, you will also learn the difference between the molecular mass and the formula mass of a compound.

Formula Mass

The **formula mass** of a substance is the sum of the atomic masses of all the atoms in one molecule or particle of that substance, in atomic mass units (amu). This terminology applies only to ionic substances, such as NaCl. For covalent molecules, such as water (H_2O), the sum of the atomic masses of all of the atoms in a molecule is called the **molecular mass**. Refer to your periodic table (See Appendix E: Periodic Table of Elements or Appendix F: Alphabetical Listing of the Elements and their Atomic Masses at the end of the course) to find the atomic masses of the elements you need to work with.

Example 1:

Find the molecular mass of water (H_2O).

One molecule of water contains 2 hydrogen atoms and one oxygen atom.

The atomic mass of hydrogen is 1.0 μ and the atomic mass of oxygen is 16.0 μ.

$$H_2O = (2 \times H) + (1 \times O) = (2 \times 1.0\,\mu) + (1 \times 16.0\,\mu) = 18.0\,\mu$$

The molecular mass for water is 18.0 μ.

Example 2:

What is the formula mass of $Ca_3(PO_4)_2$?

Remember that a subscript found outside of the parentheses means that you will have multiple ions present. A formula of $Ca_3(PO_4)_2$ means that two phosphate (PO_4^{3-}) ions are present. This means that there are two Ps and eight Os, NOT one P and eight Os.

$$Ca_3(PO_4)_2 = (3 \times Ca) + (2 \times P) + (8 \times O) =$$
$$(3 \times 40.1\,\mu) + (2 \times 31.0\,\mu) + (8 \times 16.0\,\mu)$$

$$Ca_3(PO_4)_2 = 120.3\,\mu + 62.0\,\mu + 128.0\,\mu = 310.3\,\mu$$

The formula mass for calcium phosphate is 310.3 μ.

Example 3:

What is the formula mass for $CuSO_4 \cdot 5H_2O$?

This form of copper sulfate is known as a hydrate. **Hydrates** are solids that have water molecules tightly associated with the solid crystal. The formula mass of the compound includes the associated waters. In this example, $CuSO_4 \cdot 5H_2O$ indicates there are 5 water molecules associated with each copper (II) sulfate particle.

$$CuSO_4 \cdot 5H_2O = 1\ Cu + 1\ S + 4\ O + (5\ H_2O)$$

$$CuSO_4 \cdot 5H_2O = (1 \times 63.5\,\mu) + (1 \times 32.1\,\mu) + (4 \times 16.0\,\mu) +$$
$$(5 \times (2.0\,\mu + 16.0\,\mu)) = 63.5\,\mu + 32.1\,\mu + 64.0\,\mu + 90.0\,\mu = 249.6\,\mu$$

The formula mass of $CuSO_4 \cdot 5H_2O$ is 249.6. μ.

Learning Activity 3.5

Finding Formula Mass

Find the formula mass of each of the following. Show your work.

1. NaCl
2. $Fe(C_2H_3O_2)_3$
3. $Al(NO_3)_3$
4. $KBrO_3$
5. XeF_6
6. $C_3H_5N_3O_3$ (nitroglycerin)
7. $C_{63}H_{88}N_{14}PCo$ (vitamin B_{12})
8. $MnCl_2 \cdot 4H_2O$
9. PbI_2
10. $MgCO_3$

Lesson Summary

In this lesson, you learned that the formula or molecular mass of a substance is the sum of the atomic masses of all the atoms in one molecule or particle of that substance, in atomic mass units (amu). In the next lesson, you will learn how to write and classify balanced chemical equations from written descriptions of reactions.

NOTES

Assignment 3.3

Determining Formula Mass *(10 marks)*

Find the formula mass of each of the following. Each question is worth 2 marks—one mark for showing your work, and the other mark for the correct answer. *(10 marks)*

1. K_2CO_3

2. $(NH_4)_2Cr_2O_7$

3. $FePO_4$

4. LiHS

5. $Ca(ClO_4)_2$

NOTES

Lesson 4: Balancing Chemical Equations
(2 hours)

Lesson Focus

SLO C11-3-05: Write and classify balanced chemical equations from written descriptions of reactions.
Include: polyatomic ions

Lesson Introduction

Building a home from the ground up is not a guessing game. Exact amounts of materials, such as wood, stucco, and insulation are required to make the finished product. Likewise in chemistry, you can determine how much reactant is required to make a certain amount of product. In Grade 10 Science, you may have learned that a balanced chemical equation can tell you about the quantities of reactants and products in a reaction. You may also have learned how to balance chemical equations in order to obey the Law of Conservation of Mass.

In this lesson, you will balance more complex equations and classify different types of chemical reactions.

Chemical Equations

This section provides a brief overview of what you need to know before you attempt to balance more complex chemical equations. For some, this may be a review, while for others this may be new material.

A **chemical equation** indicates the substances reacting and the substances produced in a chemical reaction. Generally, we call the substance or substances that react together the **reactants**, and we call the resulting substance or substances the **products** of a reaction. A chemical equation will usually be written with the reactants on the left side of an arrow and the products on the right side of the arrow.

REACTANTS → PRODUCTS

Perhaps you have already learned about the symbols that are unique to the language of chemistry. A chemical equation also shows the ratio in which these substances react or are produced. For example:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

You can also use words instead of symbols to describe a chemical reaction. Here is the **word equation** that describes the same chemical reaction you see above:

> "Hydrogen gas and oxygen gas react to form (or yield) water vapour."

A balanced chemical reaction provides the same information that a recipe does. In addition to symbols, a chemical reaction uses numbers to indicate the quantity of reactants used and products created. For example, the "2"s in front of both the H_2 and the H_2O are called **coefficients**. Coefficients indicate the ratio in which the substances combine or are produced in a chemical reaction. The number "1" is not written as a coefficient, so the coefficient for O_2 is 1. The balanced word equation for the above example is:

> "Two molecules of hydrogen gas react with one molecule of oxygen gas to yield two molecules of water."

Chemists show the states or phases of the reactants and products by using abbreviations in parentheses following each reactant and product. The abbreviations are as follows:

(***s***) or (***c***) means solid or crystalline

(***l***) means liquid

(***g***) means gas or vapour

(***aq***) means the substance is aqueous or dissolved in water (aqua = water)

Balanced Chemical Equations

According to the **Law of Conservation of Mass** (Matter), mass (matter) cannot be created or destroyed. During a chemical reaction then, since atoms are a form of matter, atoms cannot be created or destroyed. This means the number of each kind of atom on the reactant side of the equation must equal the number of that same kind of atom on the product side of the equation.

When you looked at the example in the previous section, you might have wondered why it was not written as

$$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$$

We call this equation *unbalanced* because the atoms are not conserved. In other words, there are not an equal number of atoms on each side of the chemical equation. While there may be two H atoms on the reactant side and two H atoms on the product side, there are two O atoms on the reactant side compared to only one O atom on the product side. The coefficients in the reaction from the previous section ensure the equation is balanced. A **balanced chemical equation** has an equal number of atoms of each element on both the reactant and product sides of the equation. The balanced chemical equation for the burning of hydrogen in oxygen is the one we saw previously:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

In this reaction, there are four H atoms on the reactant side and four H atoms on the product side. Furthermore, there are two O atoms on the reactant side and two O atoms on the product side. The Law of Conservation of Mass has been respected, and this chemical equation is balanced.

Balancing Chemical Equations

When balancing a chemical equation you *cannot change the subscripts* of the substances in the reaction. Changing subscripts changes the identity of the compounds. To balance atoms, you can *insert coefficients* rather than change subscripts. Keep the following in mind:

- Sometimes an equation is already balanced, and you won't need to change anything.
- In a balanced equation, each side of the equation will have the same number of atoms of each element.
- It is not necessary for coefficients to be the same on both sides of the equation in order for the number of atoms of each type to balance.

You may already have a system of balancing equations that works best for you. If not, or if you need a review, then you can use the following steps:

1. If it has not already been done for you, *write out the formulas of each reactant and product*. This is called the **skeleton equation**.
2. *Determine the number of atoms of each element*. Write each element in a table and keep a tally of the number of atoms of each element (see the examples below).
3. *Use coefficients to balance the equation* so that it obeys the Law of Conservation of Mass.
4. If necessary, *reduce the coefficients to the lowest whole number ratio*. Multiply fractions, if present, by the denominator to make all coefficients whole numbers.

Example 1:

Balance the equation $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$.

Step 1: *Write out the formulas of each reactant and product.*

This has already been done, so we can move directly to Step 2.

Step 2: *Determine the number of atoms of each element.*

	C_3H_8	$+ O_2$	$\rightarrow CO_2$	$+ H_2O$.
C	3		1	
H	8			2
O		2	2	1

Note: There is a total of 2 + 1 = 3 Os on the right side of the equation.

Step 3: *Use coefficients to balance the equation.* Balance the Cs and Hs first, and leave the O_2 molecule until last because it will be the easiest to balance.

	C_3H_8	$+ O_2$	$\rightarrow$ **3**CO_2	+ **4**H_2O.
C	3		1 x **3** = 3	
H	8			2 x **4**=8
O		2	2 x **3** = 6	1 x **4** = 4

There is now a total of 10 Os on the right side of the equation.

The Cs and Hs are now balanced, so we can focus on balancing the Os.

	C_3H_8 +	**5**O_2	$\rightarrow$	**3**CO_2	+	**4**H_2O.
C	3			1 x **3** = 3		
H	8					2 x **4** = 8
O		2 x **5** = 10		6		4

The balanced equation is $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$.

Step 4: *Reduce the coefficients to the lowest whole number ratio.*

The previous equation cannot be further reduced, so the answer is complete.

Example 2:

Balance the equation $Al_2(SO_4)_3 + CaCl_2 \rightarrow AlCl_3 + CaSO_4$.

Step 1: *Write out the formulas of each reactant and product.*

This has already been done, so we can move directly to Step 2.

Step 2: *Determine the number of atoms of each element.*

Since the sulfate ion remains unchanged from reactant to product, we can balance the polyatomic ion as if it were a single element, or we could balance S and O atoms separately. Note: *We can only balance the sulfate as a group if the polyatomic ion remains unchanged from reactant to product.*

	$Al_2(SO_4)_3$	$CaCl_2$	$AlCl_3$	$CaSO_4$
Al	2		1	
SO_4	3			1
Ca		1		1
Cl		2	3	

Step 3: *Use coefficients to balance the equation.* Balance the metals first. Be sure to recount other elements that are affected by the change.

	$Al_2(SO_4)_3$	$CaCl_2$	$\mathbf{2}AlCl_3$	$CaSO_4$
Al	2		1 x **2** = 2	
SO_4	3			1
Ca		1		1
Cl		2	3 x **2** = 6	

The aluminum and the calcium are balanced, but the changes have produced 6 Cls. Now we must focus on balancing the rest of the equation. Start with the Cls.

	$Al_2(SO_4)_3$	$\mathbf{3}CaCl_2$	$\mathbf{2}AlCl_3$	$CaSO_4$
Al	2		1 x **2** = 2	
SO_4	3			1
Ca		1 x **3** = 3		1
Cl		2 x **3** = 6	3 x **2** = 6	

Note: Balancing the Cls has caused an imbalance in the Cas. We must go back and balance the Cas again.

	$Al_2(SO_4)_3$ +	**3**$CaCl_2$ →	**2**$AlCl_3$ +	**3**$CaSO_4$
Al	2		1 x 2 = 2	
SO_4	3			1 x 3 = 3
Ca		1 x 3 = 3		1 x 3 = 3
Cl		2 x 3 = 6	3 x 2 = 6	

Balancing the Ca also balances the sulfate ions. The reactants and products are now balanced.

The balanced reaction is $Al_2(SO_4)_3 + 3CaCl_2 \rightarrow 2AlCl_3 + 3CaSO_4$.

Step 4: *Reduce the coefficients to the lowest whole number ratio.*

The previous equation cannot be further reduced, so the answer is complete.

If you have access to the Internet, and need more practice balancing equations, you can check out tutorials at

- www.personal.kent.edu/~cearley/ChemWrld/balance/balance.htm
- www.mpcfaculty.net/mark_bishop/balancing_equations_tutorial.htm

An interactive balancing game will provide extra practice problems at

- http://funbasedlearning.com/chemistry/chembalancer/default.htm
- http://funbasedlearning.com/chemistry/chembalancer2/default.htm
- http://funbasedlearning.com/chemistry/chemBalancer3/default.htm

Classifying Reactions

There are five general types of chemical reactions. You may have already learned about some, or all, of these reaction types. The types of reactions are **synthesis** (also called **combination**), **decomposition**, **single-replacement**, **double-replacement**, and **combustion**. In this lesson, you will put your knowledge to use and learn to recognize reaction types and ultimately predict the products of a reaction.

Synthesis Reactions

As the name suggests, synthesis reactions synthesize a product. Generally, synthesis reactions involve the reaction of two simple substances to produce a single, more complex substance.

The general form of a synthesis reaction is

$$A + B \rightarrow AB$$

A and B are reactants, usually elements or simple compounds, and AB is the product. Here are a couple of examples:

$8Fe_{(s)} + S_{8(s)} \rightarrow 8FeS_{(s)}$

$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$ (sulfuric acid)

$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$ (carbonic acid)

Note in the previous two reactions a non-metal oxide reacting with water produced an acid.

Remember that when you predict the products of a reaction, their charges must balance. In the second example above, Na^+ and Cl^- combine to form NaCl, *not* $NaCl_2$. You must always remember the ion charges when writing formulas for ionic compounds, a skill you learned in Lesson 2. In addition, you must finish your work by balancing the chemical equation.

Decomposition Reactions

Decomposition reactions are the opposite of synthesis reactions. These reactions have only one reactant producing two or more simpler substances. The decomposition of a compound usually requires the input of energy in the form of heat, electricity, or light. The general form of a decomposition reaction is

$$AB \rightarrow A + B$$

In the above reaction, AB is a complex compound while A and B are elements or simple compounds. The products will usually be in elemental or diatomic form. Remember that the elements hydrogen, oxygen, nitrogen, and the halogens all exist as diatomic molecules in their elemental form (H_2, O_2, N_2, etc.).

Here are some examples of decomposition reactions:

$2Na_2O_{(s)} \rightarrow 4Na_{(s)} + O_{2(g)}$

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

$H_2CO_{3(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$

Single Replacement Reactions

Single replacement (also called single displacement) reactions are reactions between a compound and an element. The element replaces (or displaces) one element in the compound to produce a new element and a new compound. You can identify this type of reaction by noting that both the reactants and the products consist of an element and a compound. In these reactions, metals replace metals (cations replace cations) and non-metals replace non-metals (anions replace anions). The general form of a single replacement reaction is

$$A + BX \rightarrow B + AX$$

A and B are usually metals, while X is a non-metal (negative ion). BX is often an aqueous compound (i.e., it is dissolved in water).

Single replacement reactions may also be expressed as

$$Y + AX \rightarrow X + AY$$

where X and Y represent non-metals (usually halogens), and A is a metal cation (positive ion). AX is often an aqueous compound.

A metal can replace a metal ion in an aqueous solution. For example,

$$Cu(s) + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$$

In this example, the copper and silver change positions. The copper becomes an ion and the silver becomes a solid metal. When the copper atom becomes an ion, it usually has a 2^+ charge. Remember that when ions combine, the charges must balance. Since the copper is 2^+ and the nitrate is 1^-, to make copper (II) nitrate you need two nitrate ions. Don't forget that the equation then needs to be balanced!

Here are more examples of single replacement reactions:

$$2Na_{(s)} + 2HOH_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$
$$Ca_{(s)} + 2HOH_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$
$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$
$$Cl_{2(g)} + 2NaBr_{(aq)} \rightarrow 2NaCl_{(aq)} + Br_{2(g)}$$

Note that in the final equation, the non-metals changed positions.

It is important to recognize that HOH, H_2O, and the word "water" represent the same thing. In reactions where the water molecule **dissociates** (breaks up into its ions), writing water as HOH makes it easier to see how the molecule gives a hydroxide ion to the other reactant.

Double Replacement Reactions

Double replacement (or double displacement) reactions occur when the positive ions of two compounds change positions. These types of reactions generally produce a precipitate, a gas, or a molecular compound such as water. The general form of this type of reaction is

$$AX + BY \rightarrow BX + AY$$

Here is an example of a double replacement reaction:

$$Ca(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + CaCO_{3(s)}$$

In the above reaction, the calcium and sodium ions change "partners," resulting in calcium carbonate and sodium nitrate. (It's kind of like calcium and sodium go to the school dance with their dates. They end up switching partners and dancing with each other's date.) Note that the calcium carbonate does not dissolve and precipitates out of solution.

Here is one more example of a double replacement reaction that can be written in two different ways:

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + HOH_{(l)}$

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Here again you will see water written as both HOH and H_2O. Remember that you can write the formula for water either way.

Some reactions involve a two-step process, involving two different reaction types. For example:

$Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + SO_{2(g)}$

The first step in achieving this overall reaction would be a double displacement reaction:

$Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2(SO_3)_{(aq)}$

The second step would involve the decomposition of the hydrogen sulfite:

$H_2(SO_3)_{(aq)} \rightarrow H_2O_{(l)} + SO_{2(g)}$

Combustion Reactions

A combustion reaction is a chemical change where oxygen reacts with another element or compound, often resulting in the production of heat or light energy. You can usually identify this type of reaction by looking for oxygen on the reactant side of the chemical equation. Often, the other reactant is a **hydrocarbon** (a compound composed of hydrogen and carbon) such as methane (CH_4). However, the other reactant can sometimes be another element like sulfur or magnesium, which burn in the presence of oxygen. The general equation for a combustion reaction is:

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \rightarrow xCO_2 + \left(\frac{y}{2}\right)H_2O$$

Here are some examples of combustion reactions:

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$

$C_7H_{16(l)} + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O_{(l)}$

If you have Internet access, an animation of a combustion reaction may be seen at www.marymount.k12.ny.us/marynet/stwbwk05/05flashchem/lyreaction/lyreaction.html. A YouTube video animating all five different reaction types may be observed at www.youtube.com/watch?v=tE4668aarck.

Learning Activity 3.6

Balancing Chemical Equations

1. Balance the following equations, showing your work for each example.
   a) $C_2H_4 + O_2 \rightarrow CO_2 + H_2O$
   b) $NH_3 + O_2 \rightarrow HNO_3 + H_2O$
   c) $Al(OH)_3 + Na_2SO_4 \rightarrow Al_2(SO_4)_3 + NaOH$
   d) $CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$
   e) $Al_2(SO_4)_3 + Ca(OH)_2 \rightarrow Al(OH)_3 + CaSO_4$
   f) $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$
2. Classify each of the following reactions as being synthesis, decomposition, single replacement, or double replacement. Balance any equations that are not already balanced.
   a) $Ca(OH)_{2(aq)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$
   b) $K_{(s)} + H_2O_{(l)} \rightarrow KOH_{(aq)} + H_{2(g)}$
   c) $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + Ag_{(s)}$
   d) $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$
   e) $Al(NO_3)_{3(aq)} + H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + HNO_{3(aq)}$
   f) $PbO_{2(s)} \rightarrow PbO_{(s)} + O_{2(g)}$
3. For each of the following reactions, write the balanced chemical equation and then identify the reaction type.
   a) Aqueous sodium sulfide and aqueous cadmium nitrate produce a cadmium sulfide precipitate and aqueous sodium nitrate.
   b) Solid zinc reacts with aqueous copper (II) nitrate to form elemental (solid) copper and aqueous zinc nitrate.

Lesson Summary

In this lesson, you furthered your understanding of chemical reaction types and the importance of balancing equations. You also learned how to identify the main types of chemical reactions as being synthesis, decomposition, single replacement, double replacement, or combustion.

In the next lesson, you will use your knowledge of reaction types to predict the products of a chemical reaction based on the reactants that are given.

Assignment 3.4

Classifying and Balancing Equations *(10 marks)*

1. Classify each reaction as being synthesis, decomposition, single replacement, or double replacement. Balance any equations that are not already balanced. *(1 mark for balancing, 1 mark for classification = 8 marks)*

   a) $Cl_{2(g)} + LiI_{(aq)} \rightarrow LiCl_{(aq)} + I_{2(g)}$

   ____

   ____

   b) $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$

   ____

   ____

   c) $Pb(NO_3)_{2(aq)} + NaI_{(aq)} \rightarrow PbI_{2(aq)} + NaNO_{3(aq)}$

   ____

   ____

   d) $P_2O_{5(s)} + BaO_{(s)} \rightarrow Ba_3(PO_4)_{2(s)}$

   ____

   ____

2. Solid potassium and liquid water react to form aqueous potassium hydroxide and hydrogen gas. Write the balanced chemical equation *(1 mark)* and then identify the reaction type. *(1 mark)*

   ____

   ____

Notes

Lesson 5: Chemical Reactions (1 hour)

Lesson Focus

SLO C11-3-06: Predict the products of chemical reactions, given the reactants and type of reaction.
Include: polyatomic ions

Lesson Introduction

Now that you are familiar with the five reaction types and how to balance a chemical equation, you are ready for the next step. By studying the reactants in a chemical equation, you can predict the possible reaction type and the products that could be formed.

In this lesson, you will work on recognizing the type of reaction that two reactants will most likely undergo in order to predict the products of the reaction.

Predicting the Products of a Chemical Reaction

In this lesson, you will put your knowledge of reaction types to work. When only reactants are given, you must use this information to identify the possible reaction type, and then predict the products. A helpful hint here is to determine the number of elements and/or compounds on the reactant side of the equation. For example, if there is only one reactant, you would likely suspect the reaction type to be decomposition. Once you have an idea of the possible reaction type and products, use the following steps to help you predict a balanced chemical equation for a reaction:

Step 1: *Write out the word equation (both reactants and possible products).*

Step 2: *Replace the words with formulas.*

Step 3: *Balance the final equation.*

Synthesis Reactions

In synthesis reactions, the reactants are generally two elements or compounds, while the products are a single compound. Remember that the general form of a synthesis reaction is

$$A + B \rightarrow AB$$

Example 1

Predict the product of the reaction between zinc and oxygen gas. Write out and balance the chemical equation.

This is a reaction between two elements, meaning that it is probably a synthesis reaction, where the two reactants will combine to form a new product. When predicting the product of a synthesis reaction, try to choose the simplest possible product.

Step 1: *Write out the word equation.*

zinc + oxygen gas → zinc oxide

Step 2: *Replace the words with formulas.*

For now, you do not need to indicate the states of the reactants and products. Later you will be able to predict the states of the products of some reactions.

What is the formula of the product formed in this reaction? Zinc has a charge of 2^+ and oxygen has a charge of 2^-. In Lesson 2, you learned how to determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges. The lowest common multiple of 2^+ and 2^- is 2, so only one of each ion is required. Therefore,

$Zn + O_2 \rightarrow ZnO$

Step 3: *Balance the final equation.*

$2Zn + O_2 \rightarrow 2ZnO$

Example 2

Predict the product of a reaction between sulfur dioxide and water. Write out and balance the chemical equation.

Both reactants are simple compounds. This is also a reaction between a non-metal oxide and water. From examples in Lesson 4, you should recognize this as a synthesis reaction that produces an acid.

Step 1: *Write out the word equation.*

sulfur dioxide and water → an acid

Step 2: *Replace the words with formulas.*

$SO_2 + H_2O \rightarrow H_2SO_3$

Step 3: *Balance the final equation.*

The reaction is already balanced. The product of the reaction is sulfurous acid, a source of acid rain. In this course, you will not be expected to know how to name acids.

Decomposition Reactions

Reactants in decomposition reactions are generally a single binary compound, or a compound with a polyatomic ion. The products are generally two (or more) elements or compounds, one of them usually being a gas such as carbon dioxide. The general form of a decomposition reaction is

$$AB \rightarrow A + B$$

Example 1

Predict the products of the decomposition of NaCl. Write out and balance the chemical equation.

NaCl is a binary ionic compound. The products will most likely be sodium and chlorine when it decomposes. Remember that chlorine is an element that cannot exist on its own, and must be written as a diatomic molecule, Cl_2.

Step 1: *Write out the word equation.*

sodium chloride → sodium + chlorine

Step 2: *Replace the words with formulas.*

$NaCl \rightarrow Na + Cl_2$

Step 3: *Balance the final equation.*

$2NaCl \rightarrow 2Na + Cl_2$

Single Replacement Reactions

In single replacement reactions, the reactants are usually an element and a compound. They generally produce a different element and a new compound. The general form for single replacement reactions is

$$A + BX \rightarrow B + AX$$

Example 1

Predict the product of the reaction between solid iron and aqueous copper (II) nitrate. Write out and balance the chemical equation.

This reaction involves a solid metal and an aqueous ionic compound; this is a single replacement reaction. The iron metal atom and the copper metal ion will change positions to produce copper metal and iron (III) nitrate (this formula is obtained by taking the lowest common multiple of 3^+ and 1^-, the charges of the iron and nitrate ions).

Note: Iron is an element that produces two types of ions. So iron (II) nitrate could have been a product as well. While either answer would be acceptable, we will use iron (III) in this example because it is the most common form of the iron ions. Refer to your table of ions to identify the common forms of ions that have more than one possible charge (Appendix C and Appendix D at the end of the course). The most common form of the ions will be bolded.

Step 1: *Write out the word equation.*

iron + copper (II) nitrate → copper + iron (III) nitrate

Step 2: *Replace the words with formulas.*

$$Fe_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow Cu_{(s)} + Fe(NO_3)_{3(aq)}$$

Step 3: *Balance the final equation.*

$$2Fe_{(s)} + 3Cu(NO_3)_{2(aq)} \rightarrow 3Cu_{(s)} + 2Fe(NO_3)_{3(aq)}$$

Example 2

Predict the products of the reaction between solid magnesium and aqueous sulfuric acid (H_2SO_4, hydrogen sulfate). Write out and balance the chemical equation.

This is a reaction between a metal and an aqueous compound—a single replacement reaction. The product of a metal and an acid is hydrogen gas and an ionic compound (in this case magnesium sulfate).

Step 1: *Write out the word equation.*

magnesium + sulfuric acid $\rightarrow$ hydrogen gas + magnesium sulfate

Step 2: *Replace the words with formulas.*

$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$

The formula for magnesium sulfate is found by taking the lowest common multiple of 2^+ and 2^-, the charges of the magnesium and sulfate ions.

Step 3: *Balance the final equation.*

The equation is already balanced.

Double Replacement Reactions

The reactants in double replacement reactions are two ionic compounds that exchange cations. The products are two new compounds, often with the formation of a precipitate, a gas or water molecules. The general form for the reaction is

$$AX + BY \rightarrow BX + AY$$

Example 1

Predict the product of the reaction between aqueous sodium sulfate and aqueous barium nitrate. Write out and balance the chemical equation.

This is a reaction between two aqueous compounds, hence probably a double replacement reaction. The reaction involves the ions changing "partners." The sodium will go with the nitrate and the barium will go with the sulfate. Remember that positive ions are *always* paired with negative ions, *never* positive with positive or negative with negative.

Step 1: *Write out the word equation.*

sodium sulfate$_{(aq)}$ + barium nitrate$_{(aq)}$ → sodium nitrate + barium sulfate

Step 2: *Replace the words with formulas.*

$Na_2SO_{4(aq)} + Ba(NO_3)_{2(aq)} \rightarrow NaNO_3 + BaSO_4$

The formula for sodium nitrate is found by taking the lowest common multiple of 1^+ and 1^-, the charges of the sodium and nitrate ions. The formula for barium sulfate is found by taking the lowest common multiple of 2^+ and 2^-, the charges for the barium and sulfate ions.

Step 3: *Balance the final equation.*

$Na_2SO_{4(aq)} + Ba(NO_3)_{2(aq)} \rightarrow 2NaNO_3 + BaSO_4$

For now, you will not need to predict the states of the products.

Combustion Reactions

In combustion reactions, the reactants are usually oxygen and a compound containing C, H, and O. As long as there is enough oxygen present, the products are always CO_2 (carbon dioxide gas) and H_2O (water vapour).

Example 1

Predict the products of the reaction between methane gas and oxygen gas. Write out and balance the chemical equation.

Oxygen as a reactant indicates a possible combustion reaction. Methane gas contains C and H. These reactants could form carbon dioxide and water vapour as products.

Step 1: *Write out the word equation.*

Methane and oxygen → carbon dioxide and water

Step 2: *Replace the words with formulas.*

$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Step 3: *Balance the final equation.*

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

Learning Activity 3.7

Predicting the Products of a Reaction

Predict the type of chemical reaction first, and then predict the products. Write a balanced chemical equation for each reaction.

a) sodium hydroxide$_{(aq)}$ + phosphoric acid, $H_3PO_{4(aq)} \rightarrow$

b) heating potassium carbonate$_{(s)} \rightarrow$

c) magnesium$_{(s)}$ + oxygen$_{(g)} \rightarrow$

d) chlorine$_{(g)}$ + magnesium iodide$_{(aq)} \rightarrow$

e) electrolysis of water$_{(l)} \rightarrow$

f) aluminum$_{(s)}$ + copper(II) sulfate$_{(aq)} \rightarrow$

Lesson Summary

In this lesson, you used your knowledge of chemical reactions to predict the products that could be made by a given set of reactants. You saw that the number of elements and/or compounds reacting was a good indicator of the possible reaction type and the products that might form. Extra practice predicting the products of synthesis, decomposition, single replacement, and double replacement reactions may be found at www.files.chem.vt.edu/RVGS/ACT/notes/Practice_Predicting.html. In the next lesson, you will learn about another unit of measurement in chemistry called the mole.

NOTES

Assignment 3.5

Reaction Types *(12 marks)*

1. Predict the type of reaction *(1 mark)* and then predict the products. Write the balanced chemical equation for each. *(1 mark)*

   a) ammonium sulfide$_{(aq)}$ + iron(II) nitrate$_{(aq)}$ →

   b) sodium$_{(s)}$ + chlorine$_{(g)}$ →

   c) lithium$_{(s)}$ + water$_{(l)}$ →

   d) strontium bromide$_{(aq)}$ + ammonium carbonate$_{(aq)}$ →

   e) nickel(II) carbonate$_{(s)}$ →

   f) magnesium$_{(s)}$ + hydrochloric acid, $HCl_{(aq)}$ →

NOTES

LESSON 6: THE MOLE (1 HOUR)

Lesson Focus

SLO C11-3-07: Describe the concept of the mole and its importance to measurements in chemistry.

SLO C11-3-08: Calculate the molar mass of various substances.

Lesson Introduction

Counting the eggs in your fridge is probably an easy task. Perhaps there are seven or eight. For an egg farmer, though, counting eggs one-by-one is neither reasonable nor practical. In fact, eggs have a grouping unit that you are familiar with—the dozen. An egg farmer would therefore count the number of eggs in groups of twelve, the same way you would buy them at the store. You can probably think of other grouping methods, like 52 cards being represented by one deck.

In chemistry, there is an important unit used to represent the number of particles in a substance. Since particles are extremely small, it makes sense that the mole should therefore be a large number per counting unit. In this lesson, you will learn about the mole and the number with which it is associated. You will also learn how to use this number to calculate the molar mass of different substances.

The Mole

Imagine a giant egg box that holds particles instead of eggs. Instead of there being 12 divots for eggs, there would be 6.02×10^{23} divots for particles. This special egg box represents the amount of particles found in one mole of a substance. Therefore, you would say that the box contains one mole of particles, much like the original carton held a dozen eggs.

What is a mole? In chemistry, a mole is not a small, furry, burrowing animal or a dark mark on the skin; it is just a name that describes a number. Much like an octet is a group of eight objects, a mole is 6.02×10^{23} of anything. For example:

6.02×10^{23} doughnuts is one mole of doughnuts

6.02×10^{23} pencils is one mole of pencils

6.02×10^{23} pennies is one mole of pennies

6.02×10^{23} is a very large number that is used to measure very small things, like formula units, ions, molecules, and atoms. It is 602 followed by 21 zeros or

602 000 000 000 000 000 000 000

If you were to name the number, you would say it as "six hundred and two sextillion." This number, which represents the mole, is called **Avogadro's Number** (symbolized by $\mathbf{N_A}$).

Avogadro

Most discoveries in science are named after a personality whose work was paramount in furthering our understanding of the concept. Lorenzo Romano Amadeo Carlo Avogadro (1776–1856) studied the works of scientists like Dalton and Gay-Lussac. He determined that gaseous particles of hydrogen, oxygen, nitrogen, etc. were not atoms but "molecules," and suggested that Dalton had confused the ideas of molecules and atoms.

In 1811, Avogadro published an article that stated equal volumes of gases at the same temperature and pressure contained the same number of particles. This is known as **Avogadro's Hypothesis** or **Avogadro's Principle**.

Avogadro's ideas were not immediately accepted because his proof came from an analysis of other scientists' data, and his ideas did not coincide with those of the other prominent scientists of his day.

Avogadro's work was left forgotten until 1860, when it was presented at a conference as a way to calculate the atomic masses of some elements. In 1865, Avogadro was vindicated, and his work was established as having been correct. Unfortunately, Avogadro had died several years earlier.

Avogadro's Number

Formula and molecular mass deal with individual atoms, formula units, and molecules. Chemists do not work with amounts of individual atoms or molecules because these particles are far too small to see or mass. Furthermore, balances or scales do not measure mass in terms of atomic mass units, so chemists need a practical unit that relates mass (in grams) to the number of particles present in a sample. Note that the word **particle** is a generic term meaning atoms, molecules, ions, or formula units.

The **mole (mol)** is the unit that relates the number of particles in a sample to its mass. The mole is defined as the number of particles in 12.0000 g of carbon atoms containing 6 protons and 6 neutrons each. In one mole of carbon, there are actually $6.022136736 \times 10^{23}$ atoms. We round that value to 6.02×10^{23}.

The scientific community felt badly about how Avogadro was treated and decided to honour him by naming the number of particles in a mole after him. That is why 6.02×10^{23} particles is called Avogadro's Number.

Here is an example to illustrate the size of Avogadro's Number:

> If the average loonie has a thickness of about 1.5 mm, Avogadro's Number of loonies (6.02×10^{23}) stacked one on top of the other would make a pile about 9.03×10^{17} km high! That's almost the distance across the Milky Way Galaxy!

Molar Mass

Avogadro's Number relates the number of particles to mass. By definition, one mole of carbon atoms has a mass of 12.0000 g. *If the mass of one mole of any atom is its atomic mass in grams*, then

- one mole of aluminum atoms has a mass of 27.0 g
- one mole of silver atoms has a mass of 107.9 g
- one mole of sodium atoms has a mass of 23.0 g
- one mole of iron atoms has a mass of 55.8 g

Similarly, *one mole of any compound has a mass equal to its formula mass in grams*. For example, since the formula mass of water is 18.0 amu, the mass of one mole of water molecules will be 18.0 g.

The mass of one mole of a substance is called **molar mass (MM)**. It is also referred to as **molecular weight** in some textbooks. The units for molar mass are grams per mole **(g/mol)**, so the molar mass of water is 18.0 g/mol.

Example 1

What is the molar mass of $Ca_3(PO_4)_2$?

First, you need to calculate the formula mass of calcium phosphate, which you learned how to do in Lesson 3.

$Ca_3(PO_4)_2$ = (3 x Ca) + (2 x P) + (4 x 2 x O) =
(3 x 40.1 μ) + (2 x 31.0 μ) + (8 x 16.0 μ)

$Ca_3(PO_4)_2$ = 310.3 μ

The formula mass for calcium phosphate is 310.3 μ.

Since the mass of one mole of $Ca_3(PO_4)_2$ is equal to its formula mass in grams, the molar mass of $Ca_3(PO_4)_2$ is 310.3 g/mol.

Example 2

What is the molar mass of lead(II) chloride, $PbCl_2$?

Calculate the formula mass.

$PbCl_2$ = (1 x Pb) + (2 x Cl) = (1 x 207.2 μ) + (2 x 35.5 μ) = 278.2 μ

The molar mass is the formula mass in g/mol, so the molar mass of lead(II) chloride is 278.2 g/mol.

Example 3

What is the molar mass of ammonium dichromate?

First, you must write the formula:

$(NH_4)_2Cr_2O_7$

Next, determine the formula mass for the compound:

$(NH_4)_2Cr_2O_7$ = (2 x N) + (8 x H) + (2 x Cr) + (7 x O)
$(NH_4)_2Cr_2O_7$ = (2 x 14.0 μ) + (8 x 1.0 μ) + (2 x 52.0 μ) + (7 x 16.0 μ)
$(NH_4)_2Cr_2O_7$ = 252.0 μ

Finally, convert to g/mol:

$(NH_4)_2Cr_2O_7$ = 252.0 g/mol

Learning Activity 3.8

Determining Molar Mass

1. Complete the following table.

Compound	Formula	Molar Mass
sodium hydroxide		
barium nitrate		
aluminum phosphate		
magnesium hydrogen carbonate		
lithium sulfate		
strontium phosphate		

2. Describe the relationship between Avogadro's number and one mole of any substance.

Lesson Summary

In this lesson, you learned that there is a practical way to measure the amount of particles present in a given amount of substance. This unit of measurement is called the mole and represents 6.02×10^{23} particles. You also learned how to calculate the molar mass of various substances. In the next lesson, you will take this one step further and calculate the molar volume of various gases.

NOTES

Assignment 3.6

Calculating Molar Mass *(5 marks)*

1. How can you calculate the molar mass of a compound? *(2 marks)*

   ___

   ___

2. Calculate the molar mass of $CaSO_4$, showing all the steps of your work. *(3 marks)*

   ___

   ___

   ___

   ___

NOTES

Lesson 7: Molar Volume (2 hours)

Lesson Focus

SLO C11-3-09: Calculate the volume of a given mass of gaseous substance from its density at a given temperature and pressure.
Include: molar volume calculation

Lesson Introduction

One mole of a liquid and a solid have widely varying volumes. For example, the volume of one mole of glucose (≈ 128 mL) is larger than the volume of one mole of water (18 mL). This is not the case with gases. As Avogadro postulated, the volume of one mole of any gas is predictable, as long as temperature and pressure are unchanged.

In this lesson, you will learn what molar volume is and how to calculate it for gases. You will also learn about standard temperature and pressure (STP) and the importance of measuring the molar volume of gases under these conditions.

Density

In order to successfully calculate molar volume, it is necessary to understand density. You may recall from Module 1 that density is a physical property that describes the mass per unit volume of a substance. The units for density are usually grams per mL or grams per cm^3. Density is calculated by the following formula:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

or

$$\text{density} = \frac{\text{grams}}{1\ \text{L}} \text{ or } \frac{\text{grams}}{1\ \text{mL}}$$

For gases, this formula is written as

$$\text{density} = \frac{\text{molar mass}}{\text{molar volume}}$$

When comparing densities, ice floats on water because it is less dense than liquid water. Hot air balloons rise because they are less dense than the cooler air surrounding them. Heating the air causes the particles to move farther apart, reducing both the number of particles per unit volume and the overall density.

Below is a list of several gases, along with their formulas, molar masses, and densities. From the previous module, you already know that gases expand significantly with small increases in temperature. Gases will also decrease in volume when pressure is increased. For this reason, chemists usually compare gases at common temperatures and pressures.

Table 1
Density of Gases at 25° C and 101.3 kPa (760 mm Hg or 1.0 atm) Pressure

Name	Formula	Molar gass (g/mol)	Density (g/L)
Ammonia	NH_3	17.0	0.696
Carbon Dioxide	CO_2	44.0	1.799
Carbon Monoxide	CO	28.0	1.145
Hydrogen	H_2	2.0	0.0826
Methane	CH_4	16.0	0.656
Neon	Ne	20.2	0.825
Oxygen	O_2	32.0	1.308
Propane	C_3H_8	44.0	1.802

Calculating the Volume of a Gas

One mole of any gas has an approximate volume of 22.4 L at STP. **STP (Standard Temperature and Pressure)** means 0^{o} C and either 101.3 kPa or 1 atm of pressure. The relationship 22.4 L = 1 mol at STP is a conversion factor you will use often when calculating molar volume. However, most reactions you deal with take place at room temperature (also referred to as ambient pressure). Therefore, **SATP (Standard Ambient Temperature and Pressure)** refers to a temperature of 25^{o} C and either 101.3 kPa or 1 atm of pressure. One mole of any gas has an approximate volume of 24.4 L at SATP. Look at the following examples to see how this conversion factor was derived.

Example 1

Using *Table 1: Density of Gases* from above, calculate the molar volume, in litres, of one mole of ammonia (NH_3) at 25° C and 101.3 kPa of pressure.

The molar mass of ammonia is 17.0 g/mol and its density is 0.696 g/L.

To calculate volume, rearrange the density equation for gases:

$$\text{molar volume} = \frac{\text{molar mass}}{\text{density}}$$

$$\text{molar volume} = \left(\frac{17.0\ \cancel{g}/\text{mol}}{0.696\ \cancel{g}/\text{L}}\right) = 24.4\ \text{L/mol}$$

One mole of ammonia would then have a volume of 24.4 L.

We could also multiply the molar mass by the density ratio that would cancel out the "g" units. Our choices would be

$$\frac{0.696\ \text{g}}{1\ \text{L}} \text{ or } \frac{1\ \text{L}}{0.696\ \text{g}}$$

Multiplying the second ratio by the molar mass will cancel out the "g" units:

$$\text{molar volume} = \frac{17.0\ \cancel{g}}{1\ \text{mol}} \times \left(\frac{1\ \text{L}}{0.696\ \cancel{g}}\right) = 24.4\ \text{L/mol}$$

Again, one mole of ammonia would have a volume of 24.4 L.

Example 2

Using the data from Table 1, calculate the volume (in litres) of one mole of carbon dioxide (CO_2) at 25° C and 101.3 kPa of pressure.

The molar mass of carbon dioxide is 44.0 g/mol and its density is 1.799 g/L.

To calculate volume, rearrange the density equation:

$$\text{molar volume} = \frac{\text{molar mass}}{\text{density}}$$

$$\text{molar volume} = \left(\frac{44.0\ \cancel{g}/\text{mol}}{1.799\ \cancel{g}/\text{L}}\right) = 24.45\ \text{L/mol} = 24.4\ \text{L/mol}$$

The molar volume of one mole of carbon dioxide would be 24.4 L.

Alternately, you could multiply the molar mass by the density ratio that would cancel out the "g" units. Our choices would be

$$\frac{1.799\ \text{g}}{1\ \text{L}} \text{ or } \frac{1\ \text{L}}{1.799\ \text{g}}$$

Multiplying the second ratio by the molar mass will cancel out the "g" units:

$$\text{molar volume} = \frac{44.0\ \cancel{\text{g}}}{1\ \text{mol}} \times \left(\frac{1\ \text{L}}{1.799\ \cancel{\text{g}}}\right) = 24.45\ \text{L/mol} = 24.4\ \text{L/mol}$$

Again, one mole of carbon dioxide would have a volume of 24.4 L.

Molar Volume

Avogadro was ahead of his time when he said that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. The particles may be different sizes, but they are spaced so far apart from each other that size does not really affect how many particles fit in the same area. In the previous section, you saw that the volume of one mole of a gas at the same temperature and pressure occupies the same volume. This is exactly what is stated by Avogadro's Hypothesis!

In Module 2, you learned that the volume of a gas is affected by variations in temperature and pressure. For example, a balloon will shrink in cold air, as the volume of gas will decrease with a decrease in temperature. For this reason, the volume of a gas must be measured at standard temperature and pressure (STP). The standard temperature is 0° C and the standard pressure is 101.3 kPa, 760 mm Hg, 760 torr or 1.00 atm. At standard temperature and pressure, the volume of one mole of any gas is 22.4 L. 22.4 L/mol is known as the **molar volume** of a gas at STP.

Calculating the Number of Moles of a Gas, Given Its Volume

We can use the molar volume to convert between the number of moles and the volume of a sample of gas at STP.

The number of moles present in a sample of gas can be determined by the formula:

$$\text{number of moles} = \frac{\text{sample volume}}{\text{molar volume}}$$

Or, we can multiply by the appropriate molar volume ratio that will cancel out the "L" units:

$$\frac{22.4\ \text{L}}{1\ \text{mole}} \text{ or } \frac{1\ \text{mole}}{22.4\ \text{L}}$$

Example 1

How many moles are in 50.0 L of oxygen gas at STP?

According to Avogadro's Hypothesis, the identity of the gas does not matter when calculating volume.

Using the formula

$$\text{number of moles} = \frac{\text{sample volume}}{\text{molar volume}}$$

$$\text{moles} = \left(\frac{50.0\ \cancel{\text{L}}}{22.4\ \cancel{\text{L}}/\text{mol}}\right) = 2.23\ \text{mol}$$

Alternatively, we can use the molar volume ratio from above that allows us to cancel out the "L" units:

$$\text{moles} = 50.0\ \cancel{\text{L}} \times \left(\frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}}\right) = 2.23\ \text{mol}$$

There are 2.23 moles of oxygen in 50.0 L of the gas at STP. Note that constant values, such as 22.4 L, are not used to determine significant figures.

Example 2

How many moles are there in 10.0 mL of carbon dioxide gas at STP?

The volume given is in mL. Since molar volume is in litres, we must first convert mL to litres:

$$\text{L} = 10.0 \text{ mL} \times \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.0100 \text{ L}$$

Using the formula

$$\text{number of moles} = \frac{\text{sample volume}}{\text{molar volume}}$$

$$\text{mol} = \left(\frac{0.0100 \cancel{\text{L}}}{22.4 \cancel{\text{L}}/\text{mol}} \right) = 4.46 \times 10^{-4} \text{ mol}$$

Alternatively, we can use the molar volume ratio that allows us to cancel out the "L" units:

$$\text{moles} = 0.0100 \cancel{\text{L}} \times \left(\frac{1 \text{ mol}}{22.4 \cancel{\text{L}}} \right) = 4.46 \times 10^{-4} \text{ mol}$$

There are thus 4.46×10^{-4} moles of carbon dioxide in 10.0 mL of the gas at STP.

Calculate Volume, Given Moles

If one mole of any gas at STP has a volume of 22.4 L, then 2 moles of a gas at STP will have a volume of 2 mol x 22.4 L/mol = 44.8 L. To calculate the volume from the number of moles, we can rearrange the previous formula to obtain

sample volume = number of moles x molar volume

or we can multiply the number of moles by the molar ratio that will cancel out the "mol" units.

Example 1

What is the volume, in litres, of 2.50 moles of methane, CH_4, at STP?

Once again, *the identity of the gas is unimportant* in these calculations, according to Avogadro's Hypothesis.

Using the formula

$$\text{sample volume} = \text{number of moles} \times \text{molar volume}$$

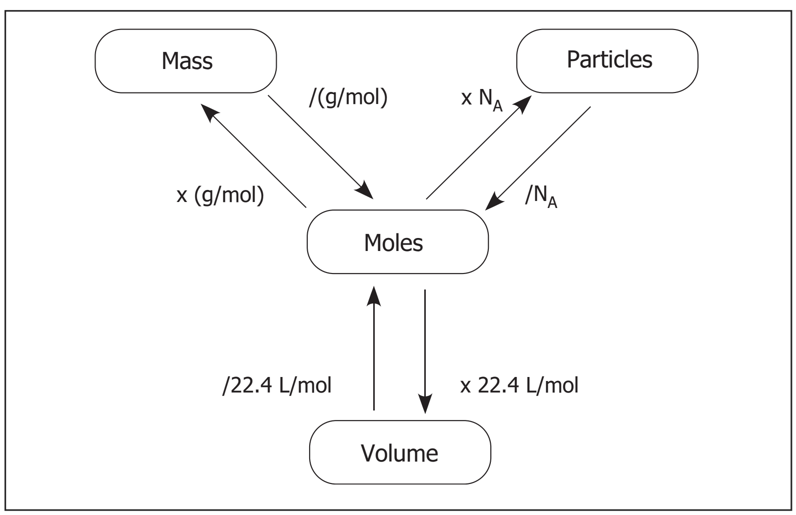
$$\text{sample volume} = 2.50\ \cancel{\text{mol}} \times 22.4\ \text{L}/\cancel{\text{mol}} = 56.0\ \text{L}$$

Using the molar volume ratio that allows us to cancel out the "mol" units

$$\text{sample volume} = 2.50\ \cancel{\text{mol}} \times \left(\frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}\right) = 56.0\ \text{L}$$

Calculating Volume, Given Mass or Particles

Sometimes when you are travelling you need to consult a road map. As you learn more conversions and types of calculations, you may need a "mole map" to help you find your way. In the next series of calculations, the mole returns to become an important step in all of your conversions. This diagram may help you organize your calculations based on the information you have and what you need to find. When using this diagram, remember that $\mathbf{N_A}$ represents Avogadro's Number (6.02×10^{23}).



It is important to note that 22.4 L/mol is the conversion factor you will use if the reaction takes place at STP conditions. If the temperature is 25° C, then you must use 24.4 L/mol as your conversion factor!

When doing conversions,

Step 1: *Convert the given quantity to moles.* If the quantity is already given in moles, then move on to step two.

Step 2: *Convert the number of moles to the desired quantity* (mass, particles, or volume).

Example 1

What is the volume of 10.0 g of oxygen gas at STP?

Step 1: *Convert the given quantity to moles.*

To convert 10.0 g to moles, we must first find the molar mass of oxygen.

Molar mass of oxygen gas = O_2 = 32.0 g/mol

Using the diagram above, you should see the first conversion is from mass → molar mass, which involves dividing the mass by the molar mass of oxygen gas. This will give you the number of moles of oxygen gas.

$$\text{mol} = \left(\frac{10.0\ \cancel{\text{g}}}{32.0\ \cancel{\text{g}}/\text{mol}}\right) = 0.3125\ \text{mol}$$

You could also multiply the mass by the molar mass ratio that would cancel out the "g" units. Our choices would be

$$\frac{32.0\ \text{g}}{1\ \text{mol}} \text{ or } \frac{1\ \text{mol}}{32.0\ \text{g}}$$

Multiplying the second ratio by 10.0 g will cancel out the "g" units.

$$10.0\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{32.0\ \cancel{\text{g}}}\right) = 0.3125\ \text{mol}$$

You are not finished the question, so don't round to the appropriate number of significant figures yet!

Step 2: *Convert the number of moles to the desired quantity.*

Using the previous diagram, you can see that to convert from moles to volume, we must multiply by the molar volume (22.4 L/mol).

$$\text{volume} = 0.3125\ \cancel{\text{mol}} \times 22.4\ \text{L}/\cancel{\text{mol}} = 7.00\ \text{L}$$

You could also convert moles to volume using the molar volume ratio that will cancel out the "mol" units. Our choices would be

$$\frac{22.4\ \text{L}}{1\ \text{mole}} \text{ or } \frac{1\ \text{mole}}{22.4\ \text{L}}$$

Multiplying the first ratio by 0.3125 mol will cancel out the "mol" units:

$$\text{volume} = 0.3125\ \cancel{\text{mol}} \times \left(\frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}\right) = 7.00\ \text{L}$$

Alternatively, you could combine both steps together like this:

$$\text{volume} = 10.0\ \cancel{\text{g}} \times \left(\frac{1\ \cancel{\text{mol}}}{32.0\ \cancel{\text{g}}}\right) \times \left(\frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}\right) = 7.00\ \text{L}$$

Does the result make sense? We should have the correct answer because all of the units cancel out and we are left with the unit of L for the final answer. This means that 10.0 g of oxygen gas, at STP, has a volume of 7.00 L.

Example 2

Calculate the volume of one million (1×10^6) molecules of natural gas, methane (CH_4), at STP. Note that *constants are not considered when determining significant figures*.

Step 1: *Convert the given quantity to moles.*

The previous diagram indicates that we can convert molecules (a type of particle) to moles by dividing by Avogadro's Number (6.02×10^{23} molecules/mol):

$$\text{mol} = \left(\frac{1 \times 10^6\ \cancel{\text{molecules}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}/\text{mol}}\right) = 1.661 \times 10^{-18}\ \text{mol}$$

You could also multiply the number of molecules by the Avogadro's Number ratio that would cancel out the "molecules" units. Our choices would be

$$\frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \text{mol}} \text{ or } \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{molecules}}$$

Multiplying by the second ratio by 1×10^6 molecules will cancel out the "molecules" units:

$$\text{mol} = 1 \times 10^6\ \cancel{\text{molecules}} \times \left(\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}}\right)$$

$$= 1.661 \times 10^{-18}\ \text{mol}$$

Do not round the value of moles yet!

Step 2: *Convert the number of moles to the desired quantity.*

The previous diagram indicates that we can convert from moles to volume by multiplying by the molar volume at STP (22.4 L/mol):

$$\text{volume} = 1.661 \times 10^{-18}\ \cancel{\text{mol}} \times 22.4\ \text{L}/\cancel{\text{mol}} = 3.72 \times 10^{-17}\ \text{L}$$

You could also multiply the number of moles by the molar volume ratio that would cancel out the "mol" units. Our choices would be

$$\frac{22.4\ \text{L}}{1\ \text{mole}} \text{ or } \frac{1\ \text{mole}}{22.4\ \text{L}}$$

The first ratio would cancel out the "mol" units:

$$\text{volume} = 1.661 \times 10^{-18}\ \cancel{\text{mol}} \times \left(\frac{22.4\ L}{1\ \cancel{\text{mol}}}\right) = 3.72 \times 10^{-17}\ \text{L}$$

Again, you could combine both steps together like this:

$$\text{volume} = \left(\frac{1 \times 10^6\ \cancel{\text{molecules}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}/\cancel{\text{mol}}}\right) \times 22.4\ \text{L}/\cancel{\text{mol}}$$

$$= 3.72 \times 10^{-17}\ \text{L}$$

or

$$\text{volume} = 1 \times 10^6\ \cancel{\text{molecules}} \times \left(\frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}}\right) \times \left(\frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}\right)$$

$$= 3.72 \times 10^{-17}\ \text{L}$$

The volume of one million molecules of natural gas, at STP, is thus 3.72×10^{-17} L. Notice that a very large number of particles can fit into a very small volume.

If you need help using the scientific notation function on your calculator, it is important that you contact your tutor/marker or a math or science teacher at your school.

Learning Activity 3.9

Working with Moles and Volume of a Gas

1. Calculate the volume, at STP, of each of the following:
   a) 2.50 moles of any gas
   b) 3.50×10^{-5} moles of a gas
   c) 3.20×10^{-2} moles of carbon dioxide gas
   d) 3.61×10^{23} atoms of helium gas
   e) 32.0 g of oxygen gas
   f) 12.0 g of fluorine gas
2. Calculate the number of moles in each of the following at STP.
   a) 5.60 L of any gas
   b) 112 L of a gas
   c) 8.96 L of fluorine gas

Calculating Mass and Number of Particles, Given the Volume of a Gas

Similarly to the previous problems, you can calculate the mass or number of particles in a given volume of gas at STP by first calculating the number of moles in the sample.

Step 1: *Convert the given volume to moles.*

Step 2: *Convert the number of moles to mass or number of particles.*

Example 1

What is the mass of 10.0 L of oxygen gas at STP?

Step 1: *Convert the given volume to moles.*

The conversion diagram shows us that to convert from volume to moles, we must divide by the molar volume:

$$\text{mol} = \left(\frac{10.0\ \cancel{\text{L}}}{22.4\ \cancel{\text{L}}/\text{mol}}\right) = 0.4464\ \text{mol}$$

Knowing that the molar volume ratio can be:

$$\frac{22.4\ \text{L}}{1\ \text{mole}} \text{ or } \frac{1\ \text{mole}}{22.4\ \text{L}}$$

We could also complete this step by choosing the ratio that will cancel out the "L" units:

$$\text{mol} = 10.0\ \cancel{\text{L}} \times \left(\frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}}\right) = 0.4464\ \text{mol}$$

Do not round the number of moles yet!

Step 2: *Convert the number of moles to mass or number of particles.*

The diagram from the previous section shows that we can convert from moles to mass by multiplying by the molar mass:

Molar mass of oxygen gas = O_2 = 32.0 g/mol

$$\text{mass} = 0.4464\ \cancel{\text{mol}} \times 32.0\ \text{g}/\cancel{\text{mol}} = 14.3\ \text{g}$$

We could also multiply by the ratio of molar mass:

$$\frac{32.0\ \text{g}}{1\ \text{mol}} \text{ or } \frac{1\ \text{mol}}{32.0\ \text{g}}$$

That will cancel out the "mol" units:

$$\text{mass} = 0.4464\ \cancel{\text{mol}} \times \left(\frac{32.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 14.3\ \text{g}$$

You can try putting the equations together in one step like this:

$$\text{mass} = \left(\frac{10.0\ \cancel{\text{L}}}{22.4\ \cancel{\text{L}}}\right) \times \left(\frac{32.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 14.3\ \text{g}$$

or

$$\text{mass} = 10.0\ \cancel{\text{L}} \times \left(\frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}}\right) \times \left(\frac{32.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 14.3\ \text{g}$$

10.0 L of oxygen gas, at STP, has a mass of 14.3 g.

Learning Activity 3.10

Working with Mass, Number of Particles, and Volume of a Gas

1. Calculate the mass of each of the following at STP.
   a) 89.6 L sulfur dioxide
   b) 1.00×10^3 L C_2H_6
   c) 10.0 L chlorine gas
2. Calculate the number of molecules in each of the following at STP.
   a) 20.0 L of carbon monoxide
   b) 5.00 L of hydrogen gas
   c) 42.0 L of water vapour

Lesson Summary

In this lesson, you learned that the molar volume of any gas at STP is about 22.4 L/mol. Since the volume of a gas changes with temperature and pressure, molar volume calculations must be performed at Standard Pressure and Temperature (STP) so that the value 22.4 L/mol may be used. Using this value, you learned how to convert between the moles, molar mass, volume, and number of particles in a sample of gas.

In the next lesson, you will go on to solve more complex problems using moles, mass, volume, and number of particles.

Assignment 3.7

Determining the Volume of a Gas *(20 marks)*

1. Calculate the volume, at STP, of each of the following. Show all your work. *(4 x 2 marks = 8 marks)*

   a) 5.31×10^{24} molecules of SO_2

   b) 25.0 g of carbon dioxide gas

continued

c) 4.50×10^{23} molecules CH_4

d) 50.0 g of ammonia gas

2. Calculate the number of moles in each of the following samples of gas at STP. Show your calculations. *(2 x 2 marks = 4 marks)*

a) 28.0 L of CO_2 gas

continued

b) 0.542 mL of neon gas

3. Calculate the mass of each of the following at STP. Show your calculations. *(2 x 2 marks = 4 marks)*

   a) 50.0 L of argon gas

   b) 12.0 L of neon gas

continued

Assignment 3.7: Determining the Volume of a Gas (continued)

4. Calculate the number of molecules in each of the following at STP. Show your calculations. *(2 x 2 marks = 4 marks)*

   a) 224 L of helium gas

   b) 5.37×10^{-4} L of ammonia

LESSON 8: INTERCONVERSIONS (3 HOURS)

Lesson Focus

SLO C11-3-10: Solve problems requiring interconversions between moles, mass, volume, and number of particles.

Lesson Introduction

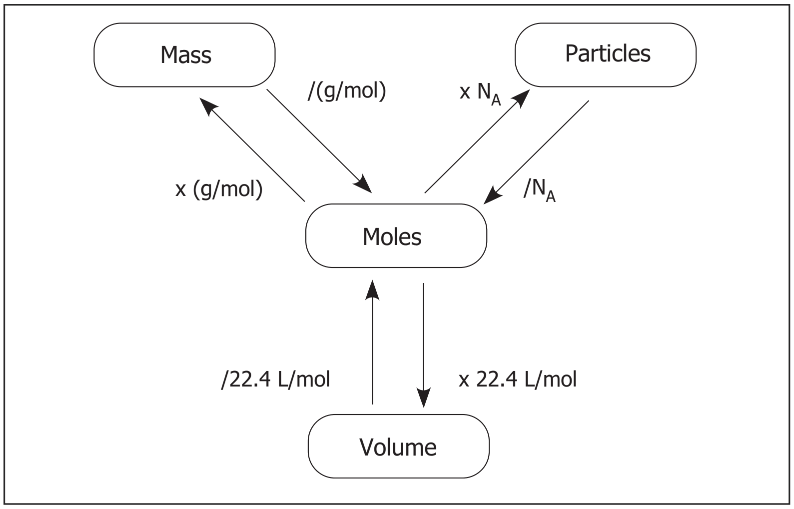
So far in Module 3, you have learned about the mole and how to calculate molar mass and molar volume. In the next series of calculations, the mole continues to be an intermediate step as you convert from one unit to another. You may find the conversion diagram can help you organize your calculations based on the information you have and what you need to find. In this lesson, you will focus on converting between moles, mass, volume, and number of particles.

The Conversion Map

In the previous lesson, you were introduced to the following conversion map to help you calculate molar volume. In this lesson, you will continue to find this road map of conversions helpful as you move between mass, volume, moles, and number of particles. On this conversion map, like the last one, $\mathbf{N_A}$ represents Avogadro's Number (6.02×10^{23}). Notice that the mole is at the centre of the map, since it is an intermediate step in all of the conversions used in this lesson. This means that you will need to convert to moles as the first step, and then convert to the unit you are seeking as a second step. We will use the same set of steps that we used in the previous lesson:

Step 1: *Convert the given quantity to moles.*

Step 2: *Convert the number of moles to the desired quantity.*



Again, it is important to note that 22.4 L/mol is the conversion factor you will use if the reaction takes place at STP. If the temperature is 25° C, then you must use 24.4 L/mol when converting between moles and volume!

Calculating Moles, Given the Mass of a Sample

The units for molar mass (MM) are g/mol (grams per mole). This represents the mass, in grams, of one mole of a substance:

$$\text{molar mass} = \frac{\text{mass in grams}}{\text{1 mole}}$$

To find the number of moles, given a sample's mass,

Step 1: *Find the molar mass.*

Step 2: *Use the molar mass to calculate the number of moles.*

Example 1

How many moles of carbon atoms are in 24.0 g of carbon?

Step 1: *Find the molar mass.*

$$\text{C} = 12.0\ \text{g/mol or } \frac{12\ \text{g}}{1\ \text{mole}}$$

Step 2: *Use the molar mass to calculate the number of moles.*

Multiplying the mass by the reciprocal of the molar mass will cancel out the "grams" and leave "moles" as the only unit. In other words, you are dividing the mass of carbon by the molar mass of carbon:

$$\text{moles C} = 24.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{12.0\ \cancel{\text{g}}} = 2.00\ \text{moles C}$$

Example 2

How many moles of water molecules are there in 81.0 g of water?

Step 1: *Find the molar mass.*

H_2O = 2H + 1O = (2 x 1.0 g/mol) + (1 x 16.0 g/mol) = 18.0 g/mol

Step 2: *Use the molar mass to calculate the number of moles.*

$$\text{moles water} = 81.0\ \cancel{g} \times \frac{1\ \text{mol}}{18.0\ \cancel{g}} = 4.50\ \text{moles water}$$

There are 4.50 moles in 81.0 g of water.

Example 3

How many moles are there in 50.0 g of lead (II) chloride, PbCl2?

Step 1: *Find the molar mass.*

$PbCl_2$ = 1Pb + 2Cl = (1 x 207.2 g/mol) + (2 x 35.5 g/mol)
= 278.2 g/mol

Step 2: *Use the molar mass to calculate the number of moles.*

$$\text{moles PbCl}_2 = 50.0\ \cancel{g} \times \frac{1\ \text{mol}}{278.2\ \cancel{g}} = 0.1797\ \text{mol} \approx 0.180\ \text{mol}$$

You must round to 3 significant figures because the number of significant figures in 50.0 g is 3.

Learning Activity 3.11

Finding the Number of Moles

Find the number of moles in each of the following:

a) 3.00 g of helium atoms
b) 6.84 g of fluorine molecules
c) 12.2 g of aluminum atoms
d) 33.5 g of iron atoms
e) 4.40 g of CO_2
f) 3.612×10^{24} atoms of oxygen
g) 1.20×10^{23} molecules of H_2O

Calculate the Mass of a Given Number of Moles

If one mole of carbon has a mass of 12.0 g, then 2 moles of carbon will have a mass of $2 \times 12.0 = 24.0$ g, 3 moles will have a mass of 3×12.0 g $= 36.0$ g, etc. If you need a certain number of moles of a substance, you can calculate the mass of the substance by multiplying the number of moles of that substance by its molar mass.

When converting from moles to mass,

Step 1: *Find the molar mass of the substance.*

Step 2: C*onvert the number of moles to mass. Multiply so that units cancel to yield grams.*

Example 1

What is the mass of 2.50 moles of gold?

Step 1: *Find the molar mass of the substance.*

Au = 197.0 g/mol

Step 2: *Convert the number of moles to mass. Multiply so that units cancel to yield grams.*

$$\text{mass Au} = 2.50\ \cancel{\text{mol}} \times \left(\frac{197.0\text{ g}}{1\ \cancel{\text{mol}}}\right) = 492.5\text{ g} \approx 492\text{ g}$$

2.50 moles of gold has a mass of about 492 g.

Example 2

What is the mass of 1.20×10^{-5} moles of carbon tetrachloride, CCl_4?

Step 1: *Find the molar mass of the substance.*

CCl_4 = (1 x 12.0 g/mol) + (4 x 35.5 g/mol) = 154.0 g/mol

Step 2: *Convert the number of moles to mass. Multiply so that units cancel to yield grams.*

$$\text{mass } CCl_4 = \left(1.20 \times 10^{-5}\ \cancel{\text{mol}}\right)\left(\frac{154.0\text{ g}}{1\ \cancel{\text{mol}}}\right) = 0.001848\text{ g}$$

$$= 1.848 \times 10^{-3}\text{ g} \approx 1.85 \times 10^{-3}\text{ g}$$

1.20×10^{-5} moles of carbon tetrachloride has a mass of about 1.85×10^{-3} g. It is more convenient to use scientific notation when values are less than 0.01 and more than 10 000.

Learning Activity 3.12

Converting Moles to Mass

Calculate the mass of each of the following:

a) 1.25 moles of NaOH

b) 0.0500 moles of KCl

c) 0.450 moles of $Mg_3(PO_4)_2$

d) 2.50×10^{-2} moles of $NaNO_3$

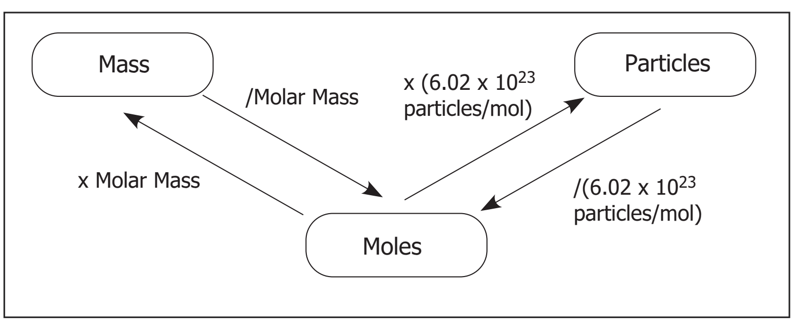
e) 4.75×10^{9} molecules of water

f) 4.515×10^{23} atoms of phosphorous

g) 1.50×10^{23} molecules of NaCl

Converting Mass to Particles

The mole allows for the conversion between the mass and number of particles in a sample. The abbreviated mole conversion map below may help to visualize the process for the conversion.



To convert from mass to a number of particles, follow the following steps:

Step 1: *Convert mass to moles.*

Step 2: *Convert the number of moles to mass. Multiply so that units cancel to yield grams.*

Example 1:

How many molecules of water are in a 10.0 g sample of water?

Step 1: *Convert mass to moles.*

To convert from mass to moles, we must know the molar mass of water.

H_2O = 18.0 g/mol

$$\text{mol} = 10.0\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{18.0\ \cancel{\text{g}}}\right) = 0.5556\ \text{mol}$$

Do not round at this point.

Step 2: *Convert the number of moles to mass. Multiply so that units cancel to yield grams.*

$$\text{molecules} = 0.5556\ \cancel{\text{mol}} \times \left(\frac{6.02 \times \text{molecules}}{1\ \cancel{\text{mol}}}\right)$$

$$= 3.34 \times 10^{23}\ \text{molecules}$$

You can also put these two equations together. Notice how all units, except "molecules," cancel out.

$$\text{molecules} = 10.0\ \cancel{\text{g}} \times \left(\frac{1\ \cancel{\text{mol}}}{18.0\ \cancel{\text{g}}}\right) \times \left(\frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}\right)$$

$$= 3.34 \times 10^{23}\ \text{molecules}$$

There are 3.34×10^{23} molecules in 10.0 g of water.

Learning Activity 3.13

Converting between Moles and Number of Particles

1. Find the number of particles in each of the following:
   a) 1.20×10^{-15} moles of zinc
   b) 4.50×10^{-7} moles of tin atoms
   c) 5.75 moles of lead
   d) 1.41 moles of water molecules
2. 75.6 g of ammonium sulfate will contain
   a) how many moles of ammonium sulfate?
   b) how many moles of nitrogen?
3. Determine the number of moles of *each element* in
   a) 2.5 moles of water
   b) 5.00 moles of NH_4NO_3
   c) 10.0 g of $CaCO_3$
   d) 50.0 g of $Ba(OH)_2$

Lesson Summary

In this lesson, you continued to use the mole as an intermediate step in converting between units of mass, volume and number of particles. By now, you should know the value of Avogadro's Number (NA) and when to use it in your conversions.

In the next lesson, you will return to chemical formulas and learn how to use percent composition or mass data to determine a molecular formula.

Assignment 3.8

Converting between Mass, Moles, and Number of Particles *(24 marks)*

1. Calculate the number of moles in each of the following. *(10 marks)*

   a) 50.0 g of water

   b) 25.0 g of silver

continued

Assignment 3.8: Converting between Mass, Moles, and Number of Particles (continued)

c) 10.0 g of sodium chloride

d) 100.0 g of sugar, $C_{11}H_{22}O_{11}$

e) 1.00 g of oxygen gas

continued

Assignment 3.8: Converting between Mass, Moles, and Number of Particles (continued)

2. Calculate the mass of each of the following. *(10 marks)*

   a) 5.00 moles of aluminum atoms

   b) 0.250 moles of nitrogen molecules

continued

Assignment 3.8: Converting between Mass, Moles, and Number of Particles (continued)

c) 2.10×10^{-3} moles of ammonia, NH_3

d) 7.25 moles of iron(III) oxide

e) 8.90×10^{-6} moles of copper(II) sulfate, $CuSO_4$

continued

Assignment 3.8: Converting between Mass, Moles, and Number of Particles (continued)

3. Find the number of particles in each of the following. *(4 marks)*
   a) 60.5 g of calcium atoms

   b) 6.31 g of HNO_3 molecules

NOTES

Lesson 9: Empirical and Molecular Formulas (1.5 hours)

Lesson Focus

SLO C11-3-11: Determine empirical and molecular formulas from percent composition or mass data.

Empirical and Molecular Formulas

When scientists analyze substances to determine their identities, they determine the elements in the substance and then calculate the relative amount of each element in that substance. Once they know relative amounts, a chemical formula can be identified. In this lesson, you will learn how to use data to determine the chemical formula of a compound.

Percent Composition

Statistics Canada is very interested in the percent composition of our Canadian population. When they publish the relative amount of Canadians who are younger than 18 years of age, the percentage of Canadians who own their own home, or the mix of males and females in our total population, they are providing examples of percent composition. In the general sense, percent composition refers to the relative size of parts that make up 100% of something.

When performing a chemical analysis to determine the identity of a compound, chemists will determine the relative amount of each element (in terms of mass) in that compound. These amounts are usually expressed as a **percent composition**, by mass. The percent composition, by mass, is determined by dividing the mass of the element in the compound by the total mass of the compound and multiplying by one hundred:

$$\text{percent composition} = \frac{\text{mass of element in compound}}{\text{total mass of compound}} \times 100\%$$

Example 1

A compound contains 22.1 g of copper, 11.2 g of sulfur, and 22.3 g of oxygen. What is the percent composition of each element in this compound?

Step 1: *Determine the total mass of the sample of the compound by adding the individual masses together.*

22.1 g + 11.2 g + 22.3 g = 55.6 g

Step 2: *Determine the percent composition of each element in the compound.*

$$\text{percent Cu} = \frac{22.1\ \cancel{g}}{55.6\ \cancel{g}} \times 100\% = 39.7\%$$

$$\text{percent S} = \frac{11.2\ \cancel{g}}{55.6\ \cancel{g}} \times 100\% = 20.1\%$$

$$\text{percent O} = \frac{22.3\ \cancel{g}}{55.6\ \cancel{g}} \times 100\% = 40.1\%$$

Percent Composition from Chemical Formulas

If you know the formula of a compound, you can also calculate the percent composition of that compound. Using the individual masses of each of the elements and the molar mass of a compound, you can go on to calculate the percent by mass of each element in one mole of that compound:

$$\text{percent composition} = \frac{\text{mass of element in 1 mol of compound}}{\text{total mass of compound}} \text{ x } 100\%$$

Example 1

Determine the percent composition of each element in Na_2CO_3.

Step 1: *Determine the number of moles of each element in one mole of the compound.*

Use the subscripts to help you determine the number of moles of each element. The number of moles of each element in one mole of the compound is equal to the subscript for that element. Thus, 1 mole of Na_2CO_3 contains 2 moles of Na, 1 mole of C, and 3 moles of O.

Step 2: *Determine the mass of each element in one mole of the compound.*

The mass of each element is the number of moles of an element multiplied by the molar mass of that element. Remember that the molar mass of an element is equal to its atomic mass expressed in grams.

$$\text{Mass of Na} = 2\ \cancel{\text{mol}} \times \frac{23.0\ \text{g}}{1\ \cancel{\text{mol}}} = 46.0\ \text{g}$$

$$\text{Mass of C} = 1\ \cancel{\text{mol}} \times \frac{12.0\ \text{g}}{1\ \cancel{\text{mol}}} = 12.0\ \text{g}$$

$$\text{Mass of O} = 3\ \cancel{\text{mol}} \times \frac{16.0\ \text{g}}{1\ \cancel{\text{mol}}} = 48.0\ \text{g}$$

Step 3: *Find the fraction of each element in the compound and convert the value to a percentage.*

To find the fraction, divide the mass of each element in one mole by the mass of one mole (molar mass) of the compound.

molar mass of Na_2CO_3 = (2 x 23.0 g/mol) + (1 x 12.0 g/mol) + (3 x 16.0 g/mol) = 106.0 g/mol

$$\text{percent Na} = \frac{46.0\ \cancel{\text{g}}}{106.0\ \cancel{\text{g}}} \times 100\% = 43.4\%$$

$$\text{percent C} = \frac{12.0\ \cancel{\text{g}}}{106.0\ \cancel{\text{g}}} \times 100\% = 11.3\%$$

$$\text{percent O} = \frac{48.0\ \cancel{\text{g}}}{106.0\ \cancel{\text{g}}} \times 100\% = 45.3\%$$

One way to check if your calculations are correct is to add up the percent compositions. Their sum should be 100%, or at least close to 100%.

43.4% + 11.3% + 45.3% = 100%

Example 2

Determine the percent, by mass, of water in the hydrate $MgSO_4 \cdot 9H_2O$ (magnesium sulfate nonahydrate).

Remember that hydrates are ionic compounds that have a specific number of water molecules loosely attached to them. To determine the percent of water in a hydrate, consider the compound and the water as two separate entities and treat the $MgSO_4$ and H_2O separately.

Step 1: *Determine the number of moles of water in one mole of the compound.*

For every mole of magnesium sulfate, there are 9 water molecules

Step 2: *Determine the mass of each water molecule in one mole of the compound.*

H_2O = (2 x 1.0 g/mol) + (1 x 16.0 g/mol) = 18.0 g/mol

$$\text{mass of } H_2O = 9 \cancel{\text{ mol}} \times \frac{18.0 \text{ g}}{1 \cancel{\text{ mol}}} = 162.0 \text{ g}$$

Step 3: *Find the fraction of each element in the compound and convert the value to a percentage.*

To find the fraction, divide the mass of water in one mole by the molar mass of the compound.

$MgSO_4 \cdot 9H_2O$ = (1 x 24.3 g/mol) + (1 x 32.1 g/mol) + (13 x 16.0 g/mol) + (18 x 1.0 g/mol) = 282.4 g/mol

$$\text{percent } H_2O = \frac{162.0 \cancel{\text{ g}}}{282.4 \cancel{\text{ g}}} \times 100\% = 57.365\% \approx 57.4\%$$

Empirical Formula

A cheeseburger contains one patty of beef and one slice of cheese. A double cheeseburger contains two patties of beef and two slices of cheese. In chemistry, you can also use basic ratios like these to determine the formula for a compound. Multiplying by the basic ratio can sometimes be used to find the formulas for other compounds as well. For example, using the above ratio, you could easily determine the number of beef patties and cheese slices in a triple cheeseburger. Once chemists have analyzed a compound and determined its percent composition, they can use this information to determine the compound's formula.

A compound can have two formulas, a **molecular formula** and an **empirical formula**. The basic ratio, called the empirical formula, is the formula of a compound having the smallest whole number ratio of the moles of each element. The empirical formula of a compound must be obtained from experimental data.

The molecular formula of a compound is the actual formula of the compound. This may or may not be the same as the empirical formula. The molecular formula contains the actual number of atoms of each element in one unit of that compound. The subscripts of the molecular formula are always a whole number multiple of the subscripts of the empirical formula.

Empirical Formula	Molecular Formula
CH_2	C_2H_4, C_7H_{14}, $C_{20}H_{40}$
HO	H_2O_2 (hydrogen peroxide)
CH_2O	$C_6H_{12}O_6$ (glucose)
Na_2CO_3	Na_2CO_3

In the examples below, you will notice that the empirical formula is the same for all three compounds. This is because, in each compound, the lowest ratio of carbon to hydrogen to oxygen is the same.

Compound	Empirical Formula	Multiplier	Molecular Formula
Formaldehyde	CH_2O	1	CH_2O
Acetic acid (vinegar)	CH_2O	2	$C_2H_4O_2$
Glucose	CH_2O	6	$C_6H_{12}O_6$

To determine the empirical formula of a substance, you need to know the elements in the compound and the percent composition of each. In some cases, this information will already be provided to you. If not, follow the steps outlined below:

1. *Assume that the total mass of the sample is 100 g.* This means the mass of each element in 100 g of the compound is equal to its percent composition.
2. *Determine the number of moles of each element in 100 g of the compound.*
3. *Calculate the lowest whole number ratio by dividing each mole value by the smallest mole value.*

Example 1

Calculate the empirical formula of a compound whose percent composition is 58.8% carbon, 9.80% hydrogen, and 31.4% oxygen.

Step 1: *Assume that the total mass of the sample is 100 g.*

That means that in a 100 g sample, there will be 58.8 g of carbon, 9.80 g of hydrogen, and 31.4 g of oxygen.

Step 2: *Determine the number of moles of each element in the 100 g sample*, dividing the mass of each element in the sample by the molar mass of that element.

$$\text{moles C} = 58.8\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{12.0\ \cancel{\text{g}}} = 4.90\ \text{mol}$$

$$\text{moles H} = 9.80\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{1.01\ \cancel{\text{g}}} = 9.80\ \text{mol}$$

$$\text{moles O} = 31.4\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{16.0\ \cancel{\text{g}}} = 1.96\ \text{mol}$$

Step 3: *Calculate the lowest whole number ratio by dividing each mole value by the smallest mole value*. This will give 1 mol for the element with the smallest number of moles. Remember that the empirical formula can only display *whole numbers* so we will not be concerning ourselves with significant figures here.

The lowest mole value is 1.96 moles of oxygen. Divide each number of moles by 1.96 moles.

$$C = \frac{4.90\ \cancel{\text{mol}}}{1.96\ \cancel{\text{mol}}} = 2.5$$ (this value is not close enough to a whole number to round)

$$H = \frac{9.80\ \cancel{\text{mol}}}{1.96\ \cancel{\text{mol}}} = 5$$

$$O = \frac{1.96\ \cancel{\text{mol}}}{1.96\ \cancel{\text{mol}}} = 1$$

Using this information, you can now determine the empirical formula of the compound to be $C_{2.5}H_5O_1$.

Since this is not the lowest *whole number* ratio, you must multiply each subscript by a value that will give the lowest whole number ratio (in this case it is 2):

$2 \times (C_{2.5}H_5O_1) = C_{(2.5 \times 2)}H_{(5 \times 2)}O_{(1 \times 2)} = C_5H_{10}O_2$

The empirical formula for the compound is $C_5H_{10}O_2$.

Molecular Formulas

The molecular formula of a compound is either the same as the empirical formula, or it is a whole-number multiple of the empirical formula. To find the molecular formula, you need to determine if the empirical formula is the same as the molecular formula or if it is some whole-number multiple of the empirical formula. Once you have determined the empirical formula of a compound, you need to calculate its molar mass to help you figure out the molecular formula. You can find this information by dividing the molecular formula by the molar mass of the empirical formula.

Example 1

Determine the molecular formula of a compound if its empirical formula is $C_5H_{10}O_2$ and its molar mass is 102.0 g/mol.

Step 1: *Determine the mass of one mole of the empirical formula.*

$C_5H_{10}O_2$ = (5 x 12.0 g/mol) + (10 x 1.0 g/mol) + (2 x 16.0 g/mol) = 102.0 g/mol

Step 2: *Divide the molar mass of the molecular formula by the molar mass of the empirical formula to obtain a whole number.*

$$\begin{aligned} \text{multiple } (\times) &= \frac{\text{molecular formula mass}}{\text{empirical formula mass}} \\ &= \frac{102.0 \text{ g/mol}}{102.0 \text{ g/mol}} \\ &= 1 \end{aligned}$$

The molecular formula is $(C_5H_{10}O_2)_x$ and the multiple is 1. $(C_5H_{10}O_2)_1$ = $C_5H_{10}O_2$. This means that the molecular formula is the same as the empirical formula.

Example 2

A compound was analyzed and found to contain 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen. If the molecular mass was determined to be 180.0 g/mol, what is the molecular formula of the compound? You must determine the empirical formula before you can determine the molecular formula.

Step 1: *Assume that the total mass of the sample is 100 g.*

In a 100 g sample, there will be 40.0 g of carbon, 6.7 g of hydrogen, and 53.3 g of oxygen.

Step 2: *Determine the number of moles of each element in the 100 g sample.*

$$40.0\ \cancel{g} \times \frac{1\ \text{mol}}{12.0\ \cancel{g}} = 3.33\ \text{mol}$$

$$6.7\ \cancel{g} \times \frac{1\ \text{mol}}{1.0\ \cancel{g}} = 6.7\ \text{mol}$$

$$53.3\ \cancel{g} \times \frac{1\ \text{mol}}{16.0\ \cancel{g}} = 3.33\ \text{mol}$$

Step 3: *Calculate the lowest whole number ratio by dividing each mole value by the smallest mole value.*

The lowest mole value is 3.33 moles of carbon and oxygen. Divide each number of moles by 3.33 moles to obtain a whole number.

$$\text{C} = \frac{3.33\ \cancel{\text{mol}}}{3.33\ \cancel{\text{mol}}} = 1$$

$$\text{H} = \frac{6.7\ \cancel{\text{mol}}}{3.33\ \cancel{\text{mol}}} = 2$$

$$\text{O} = \frac{3.33\ \cancel{\text{mol}}}{3.33\ \cancel{\text{mol}}} = 1$$

The empirical formula appears to be CH_2O.

Step 4: *Determine the mass of one mole of the empirical formula.*

CH_2O = (1 x 12.0 g/mol) + (2 x 1.0 g/mol) + (1 x 16.0 g/mol) = 30.0 g/mol

Step 5: *Divide the molar mass of the molecular formula by the molar mass of the empirical formula to obtain a whole number.*

$$\text{multiple } (\times) = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

$$= \frac{180.0 \ \cancel{\text{g/mol}}}{30.0 \ \cancel{\text{g/mol}}}$$

$$= 6$$

The molecular formula is $(CH_2O)_x$ and the multiple is 6. This means that the molecular formula is $(CH_2O)_6 = C_6H_{12}O_6$.

Learning Activity 3.14

Percent Composition, Empirical Formula, and Molecular Formula

1. Determine the percent composition of each element in the following.
   a) H_2SO_4
   b) $Fe(C_2H_3O_2)_3$
   c) $C_{12}H_{22}O_{11}$
2. What are the empirical formulas of the following compounds?
   a) $C_2H_4(OH)_2$
   b) $H_2S_2O_8$
   c) C_4H_{10}
   d) C_2H_5OH
3. Determine the empirical formula for each of the following.
   a) 0.104 mol potassium, 0.052 mol carbon, and 0.156 mol oxygen
   b) 35.98% aluminum and 64.02% sulfur
   c) 22.9% sodium, 21.5% boron, and 55.7% oxygen
   d) 21.7% carbon, 9.6% oxygen, and 68.7% fluorine

continued

Learning Activity 3.14 (continued)

4. Given the empirical formula and molar mass, what is the molecular formula of each of the following compounds?
   a) HCO_2, molar mass = 90.0 g/mol
   b) C_2H_4O, molar mass = 88.0 g/mol
   c) Na_2SiO_3, molar mass = 732.6 g/mol
5. Monosodium glutamate (MSG) is a flavour enhancer in certain foods that contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na. If MSG has a molar mass of 169.0 g/mol, what is its molecular formula?
6. Ibuprofen is a headache remedy with a molar mass about 206.0 g/mol. It contains 75.69% C, 8.80% H, and 15.51% O by mass. What is its molecular formula?

Lesson Summary

In this lesson, you learned how to determine the empirical formula, the smallest whole-number ratio of atoms in a compound. You also learned about the molecular formula of a compound, which can be the same as the compound's empirical formula or a simple whole-number multiple of the empirical formula. These concepts will help you solve stoichiometric problems in the coming lessons.

Assignment 3.9

Determining Empirical and Molecular Formulas from Percent Composition *(16 marks)*

1. Determine the percent composition of each element in aspartame ($C_{14}H_{18}N_2O_5$). *(4 marks)*

continued

Assignment 3.9: Determining Empirical and Molecular Formulas from Percent Composition (continued)

2. A sample of Freon, a gas once used as a propellant in aerosol cans, was found to contain 0.423 g of C, 2.50 g Cl, and 1.34 g F. What is the empirical formula of this compound? *(5 marks)*

3. Epinephrine, also known as adrenaline, is a hormone secreted into the bloodstream in times of danger or stress. It contains 59.0% C, 7.1% H, 26.2% O, and 7.7% N by mass and has a molecular mass of about 183.0 g/mol. What is its molecular formula? *(7 marks)*

MODULE 3 SUMMARY

Congratulations! You have reached the end of Module 3.

Submitting Your Assignments

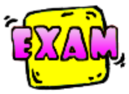
It is now time for you to submit Assignments 3.1 to 3.9 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 3 assignments and organize your material in the following order:

- Module 3 Cover Sheet (found at the end of the course Introduction)
- Assignment 3.1: Working with Isotopes
- Assignment 3.2: Working with Chemical Compounds
- Assignment 3.3: Determining Formula Mass
- Assignment 3.4: Classifying and Balancing Equations
- Assignment 3.5: Reaction Types
- Assignment 3.6: Calculating Molar Mass
- Assignment 3.7: Determining the Volume of a Gas
- Assignment 3.8: Converting between Mass, Moles, and Number of Particles
- Assignment 3.9: Determining Empirical and Molecular Formulas from Percent Composition

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

Midterm Examination



Congratulations, you have finished Module 3 in the course. The midterm examination is out of 100 marks and worth 20% of your final mark. In order to do well on this examination, you should review all of your learning activities and assignments from Modules 1 to 3.

You will complete this examination while being supervised by a proctor. You should already have made arrangements to have the examination sent to the proctor from the Distance Learning Unit. If you have not yet made arrangements to write it, then do so now. The instructions for doing so are provided in the Introduction to this module.

You will need to bring the following items to the examination: some pens and/or pencils (2 or 3 of each), an eraser or correction fluid, some blank paper, a ruler, and a scientific or graphing calculator.

A maximum of 2.5 hours is available to complete your midterm examination. When you have completed it, the proctor will then forward it for assessment. Good luck!

Grade 11 Chemistry (30S)

Module 3: Chemical Reactions

Learning Activity Answer Keys

MODULE 3: CHEMICAL REACTIONS

Learning Activity 3.1: What Is an Isotope?

1. Identify the numbers of protons, neutrons, and electrons in a neutral atom of each of the following:

   a) $^{235}_{92}\text{U}$

   b) $^{226}_{88}\text{Ra}$

   Answer:

   a) Atomic number = 92, mass number = 235
      Protons = electrons = 92
      Neutrons = 235 – 92 = 143

   b) Atomic number = 88, mass number = 226
      Protons = electrons = 88
      Neutrons = 226 – 88 = 138

2. Complete the following table to calculate the average atomic mass of each element.

Element	Symbol	Mass Number	Mass (μ)	Relative Abundance (%)	Average Atomic Mass (μ)
Carbon	C–12	12	12.000	98.98	
	C–13	13	13.003	1.11	
Silicon	Si–28	28	27.977	92.21	
	Si–29	29	28.976	4.70	
	Si–30	30	29.974	3.09	

   Answer:

$$\text{carbon: } \left(\frac{98.98\%}{100\%}\right)12.000\,\mu + \left(\frac{1.11\%}{100\%}\right)13.003\,\mu = 12.02\,\mu$$

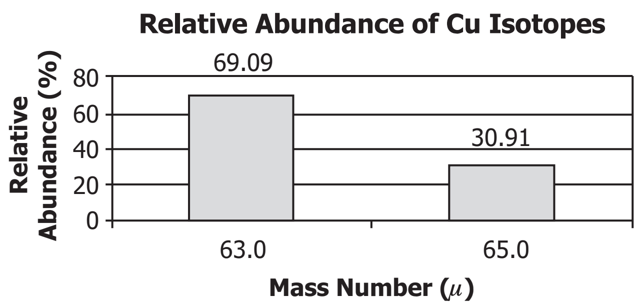
$$\text{silicon: } \left(\frac{92.21\%}{100\%}\right)27.977\,\mu + \left(\frac{4.70\%}{100\%}\right)28.976\,\mu + \left(\frac{3.09\%}{100\%}\right)29.974\,\mu = 28.08\,\mu$$

3. Define the term isotope. Explain how an element's atomic mass is related to the abundances of its different isotopes.

   Answer:

   Isotopes are atoms of the same element that differ in mass because they have different numbers of neutrons. The average mass is a weighted average of all isotopes, based on each isotope's relative abundance. The average atomic mass tends to be closest to the mass of the isotope with the greatest abundance.

4. Using the graph below, calculate the average atomic mass of copper.



   Answer:

   $$\text{Copper: } \left(\frac{69.09\%}{100\%}\right)63.0\,\mu + \left(\frac{30.91\%}{100\%}\right)65.0\,\mu = 63.6\,\mu$$

5. There are three isotopes of silicon, having the mass numbers of 28, 29, and 30. The atomic mass of silicon is 28.068 amu. Which of the isotopes is the most abundant? Which would be the least abundant?

   Answer:

   The most abundant isotope would be silicon-28, while the least abundant would be silicon-30. The average atomic mass tends to be closest to the mass of the isotope with the greatest abundance.

Learning Activity 3.2: Naming and Writing Formulas for Ionic Compounds

1. Name the following binary ionic compounds:

	Answers:
a) NiI_2	nickel (II) iodide
b) MgS	magnesium sulfide
c) K_3N	potassium nitride
d) $FeBr_2$	iron (II) bromide
e) $CaCl_2$	calcium chloride

2. Write formulas for the following binary ionic compounds:

	Answers:
a) aluminum oxide	Al_2O_3
b) tin (IV) sulfide	SnS_2
c) barium nitride	Ba_3N_2
d) chromium (III) chloride	$CrCl_3$
e) magnesium oxide	MgO

Learning Activity 3.3: Writing Formulas for Polyatomic Ions

1. Write the correct formulas for these polyatomic compounds:

	Answers:
a) Sodium hypochlorite	$NaClO$
b) Strontium nitrate	$Sr(NO_3)_2$
c) Cobalt (III) sulfite	$Co_2(SO_3)_3$
d) Copper (II) phosphate	$Cu_3(PO_4)_2$
e) Zinc hydroxide	$Zn(OH)_2$

2. Name the following polyatomic ionic compounds:

	Answers:
a) $K_2Cr_2O_7$	potassium dichromate
b) $Ni(OH)_2$	nickel (II) hydroxide
c) $AgNO_3$	silver nitrate
d) Na_3PO_4	sodium phosphate
e) $Al_2(SiO_3)_3$	aluminum silicate

Learning Activity 3.4: Naming and Writing Formulas for Covalent Compounds

1. Name the following covalent compounds:

	Answers:
a) NI_3	nitrogen triiodide
b) CO	carbon monoxide
c) SF_6	sulfur hexafluoride
d) O_2	oxygen
e) I_2	iodine

2. Write the formulas for the following covalent compounds:

	Answers:
a) Silicon tetrafluoride	SiF_4
b) Diphosphorous pentaoxide	P_2O_5
c) Tetraarsenic decaoxide	As_4O_{10}
d) Hydrogen	H_2
e) Chlorine	Cl_2

Learning Activity 3.5: Finding Formula Mass

Find the formula mass of each of the following. Show your work.

1. NaCl
2. $Fe(C_2H_3)_2)_3$
3. $Al(NO_3)_3$
4. $KBrO_3$
5. XeF_6
6. $C_3H_5N_3O_3$ (nitroglycerin)
7. $C_{63}H_{88}N_{14}PCo$ (vitamin B_{12})
8. $MnCl_2 \cdot 4H_2O$
9. PbI_2
10. $MgCO_3$

Answers:

1. $(1 \times 23.0\ \mu) + (1 \times 35.5\ \mu) = 58.5\ \mu$
2. $(1 \times 55.8\ \mu) + (6 \times 12.0\ \mu) + (9 \times 1.0\ \mu) + (6 \times 16.0\ \mu) = 232.8\ \mu$
3. $(1 \times 27.0\ \mu) + (3 \times 14.0\ \mu) + (9 \times 16.0\ \mu) = 213.0\ \mu$
4. $(1 \times 39.1\ \mu) + (1 \times 79.9\ \mu) + (3 \times 16.0\ \mu) = 167.0\ \mu$
5. $(1 \times 131.3\ \mu) + (6 \times 19.0\ \mu) = 245.3\ \mu$
6. $(3 \times 12.0\ \mu) + (5 \times 1.0\ \mu) + (3 \times 14.0\ \mu) + (3 \times 16.0\ \mu) = 131.0\ \mu$
7. $(63 \times 12.0\ \mu) + (88 \times 1.0\ \mu) + (14 \times 14.0\ \mu) + (14 \times 16.0\ \mu) + (1 \times 31.0\ \mu) +$ $(1 \times 58.9\ \mu) = 1353.9\ \mu$
8. $(1 \times 54.9\ \mu) + (2 \times 35.5\ \mu) + 4\ ((2 \times 1.0\ \mu) + (1 \times 16.0\ \mu)) = 197.9\ \mu$
9. $(1 \times 207.2\ \mu) + (2 \times 126.9\ \mu) = 461.0\ \mu$
10. $(1 \times 24.3\ \mu) + (1 \times 12.0\ \mu) + (3 \times 16.0\ \mu) = 84.3\ \mu$

Learning Activity 3.6: Balancing Chemical Equations

1. Balance the following equations, showing your work for each example.

   a) $C_2H_4 + O_2 \rightarrow CO_2 + H_2O$

   Answer:

   $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

   b) $NH_3 + O_2 \rightarrow HNO_3 + H_2O$

   Answer:

   $NH_3 + 2O_2 \rightarrow HNO_3 + H_2O$

   c) $Al(OH)_3 + Na_2SO_4 \rightarrow Al_2(SO_4)_3 + NaOH$

   Answer:

   $2Al(OH)_3 + 3Na_2SO_4 \rightarrow Al_2(SO_4)_3 + 6NaOH$

   d) $CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$

   Answer:

   $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

   e) $Al_2(SO_4)_3 + Ca(OH)_2 \rightarrow Al(OH)_3 + CaSO_4$

   Answer:

   $Al_2(SO_4)_3 + 3Ca(OH)_2 \rightarrow 2Al(OH)_3 + 3CaSO_4$

   f) $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

   Answer:

   $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

2. Classify each of the following reactions as being synthesis, decomposition, single replacement or double replacement. Balance any equations that are not already balanced.

   a) $Ca(OH)_{2(aq)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$

   Answer:

   $Ca(OH)_{2(aq)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$

   double replacement

   b) $K_{(s)} + H_2O_{(l)} \rightarrow KOH_{(aq)} + H_{2(g)}$

   Answer:

   $2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

   single replacement

c) $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + Ag_{(s)}$

Answer:

$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

single replacement

d) $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$

Answer:

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$

synthesis

e) $Al(NO_3)_{3(aq)} + H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 6HNO_{3(aq)}$

Answer:

$2Al(NO_3)_{3(aq)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + HNO_{3(aq)}$

double replacement

f) $PbO_{2(s)} \rightarrow PbO_{(s)} + O_{2(g)}$

Answer:

$2PbO_{2(s)} \rightarrow 2PbO_{(s)} + O_{2(g)}$

decomposition

3. For each of the following reactions, write the balanced chemical equation and then identify the reaction type.

a) Aqueous sodium sulfide and aqueous cadmium nitrate produce a cadmium sulfide precipitate and aqueous sodium nitrate.

Answer:

$Na_2S_{(aq)} + Cd(NO_3)_{2(aq)} \rightarrow CdS_{(s)} + 2NaNO_{3(aq)}$

double replacement

b) Solid zinc reacts with aqueous copper (II) nitrate to form elemental (solid) copper and aqueous zinc nitrate.

Answer:

$Zn_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow Cu_{(s)} + Zn(NO_3)_{2(aq)}$

single replacement

Learning Activity 3.7: Predicting the Products of a Reaction

Predict the type of chemical reaction first, and then predict the products. Write a balanced chemical equation for each reaction.

a) sodium hydroxide$_{(aq)}$ + phosphoric acid, $H_3PO_{4(aq)}$ →

Answer:

Double replacement

$3NaOH_{(aq)} + H_3PO_{4(aq)} \rightarrow Na_3PO_4 + 3H_2O$

b) heating potassium carbonate$_{(s)}$ →

Answer:

Decomposition

$K_2CO_{3(s)} \rightarrow K_2O + CO_2$

c) magnesium$_{(s)}$ + oxygen$_{(g)}$ →

Answer:

Synthesis

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO$

d) chlorine$_{(g)}$ + magnesium iodide$_{(aq)}$ →

Answer:

Single replacement

$Cl_{2(g)} + MgI_{2(aq)} \rightarrow MgCl_2 + I_2$

e) electrolysis of water$_{(l)}$ →

Answer:

Decomposition

$2H_2O_{(l)} \rightarrow 2H_2 + O_2$

f) aluminum$_{(s)}$ + copper(II) sulfate$_{(aq)}$ →

Answer:

Single replacement

$2Al_{(s)} + 3CuSO_{4(aq)} \rightarrow 3Cu + Al_2(SO_4)_3$

Learning Activity 3.8: Determining Molar Mass

1. Complete the following table.

   Answers:

Compound	Formula	Molar Mass
sodium hydroxide	$NaOH$	40.0 g/mol
barium nitrate	$Ba(NO_3)_2$	261.3 g/mol
aluminum phosphate	$AlPO_4$	122.0 g/mol
magnesium hydrogen carbonate	$Mg(HCO_3)_2$	146.3 g/mol
lithium sulfate	Li_2SO_4	109.9 g/mol
strontium phosphate	$Sr_3(PO_4)_2$	452.8 g/mol

2. Describe the relationship between Avogadro's number and one mole of any substance.

   Answer:

   One mole of any substance contains Avogadro's number of particles (6.02×10^{23}).

Learning Activity 3.9: Working with Moles and Volume of a Gas

1. Calculate the volume, at STP, of each of the following:

   a) 2.50 moles of any gas

   Answer:

$$\text{volume} = 2.50\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 56.0\ \text{L}$$

   b) 3.50×10^{-5} moles of a gas

   Answer:

$$\text{volume} = 3.50 \times 10^{-5}\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 7.84 \times 10^{-4}\ \text{L}$$

c) 3.20 x 10^{-2} moles of carbon dioxide gas

Answer:

$$\text{volume} = 3.20 \times 10^{-2}\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 0.717\ \text{L}$$

d) 3.61 x 10^{23} atoms of helium gas

Answer:

$$\text{volume} = 3.61 \times 10^{23}\ \cancel{\text{atoms}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{atoms}}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 13.4\ \text{L}$$

e) 32.0 g of oxygen gas

Answer:

$$O_2 = 32.0\ \text{g/mol}$$

$$\text{volume} = 32.0\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{32.0\ \cancel{\text{g}}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 22.4\ \text{L}$$

f) 12.0 g of fluorine gas

Answer:

$$F_2 = 38.0\ \text{g/mol}$$

$$\text{volume} = 12.0\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{38.0\ \cancel{\text{g}}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 7.07\ \text{L}$$

2. Calculate the number of moles in each of the following at STP.

a) 5.60 L of any gas

Answer:

$$\text{moles} = 5.60\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}} = 0.250\ \text{mol}$$

b) 112 L of a gas

Answer:

$$\text{moles} = 112\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}} = 5.00\ \text{mol}$$

c) 8.96 L of fluorine gas

Answer:

$$\text{moles} = 8.96\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}} = 0.400\ \text{mol}$$

Learning Activity 3.10: Working with Mass, Number of Particles, and Volume of a Gas

1. Calculate the mass of each of the following at STP.
   a) 89.6 L sulfur dioxide

      Answer:

      SO_2 = 64.1 g/mol

      $$\text{mass} = 89.6\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{64.1\ \text{g}}{1\ \cancel{\text{mol}}} = 256\ \text{g}\ \text{SO}_2$$

   b) 1.00 x 10^3 L C_2H_6

      Answer:

      C_2H_6 = 30.0 g/mol

      $$\text{mass} = 1.00 \times 10^3\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{30.0\ \text{g}}{1\ \cancel{\text{mol}}} = 1339\ \text{g}\ \ \text{C}_2\text{H}_6 \approx 1340\ \text{g}$$

   c) 10.0 L chlorine gas

      Answer:

      Cl_2 = 71.0 g/mol

      $$\text{mass} = 10.0\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{71.0\ \text{g}}{1\ \cancel{\text{mol}}} = 31.7\ \text{g}\ \text{Cl}_2$$

2. Calculate the number of molecules in each of the following at STP.
   a) 20.0 L of carbon monoxide

      Answer:

      $$\text{molecules} = 20.0\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

      $$= 5.38 \times 10^{23}\ \text{molecules}$$

   b) 5.00 L of hydrogen gas

      Answer:

      $$\text{molecules} = 5.00\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

      $$= 1.34 \times 10^{23}\ \text{molecules}$$

c) 42.0 L of water vapour

Answer:

$$\text{molecules} = 42.0\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$= 1.13 \times 10^{24}\ \text{molecules}$$

Learning Activity 3.11: Finding the Number of Moles

Find the number of moles in each of the following:

a) 3.00 g of helium atoms

Answer:

$$\text{moles} = 3.00\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{4.0\ \cancel{\text{g}}} = 0.75\ \text{mol He}$$

b) 6.84 g of fluorine molecules

Answer:

$$\text{moles} = 6.84\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{38.0\ \cancel{\text{g}}} = 0.180\ \text{mol}\ F_2$$

c) 12.2 g of aluminum atoms

Answer:

$$\text{moles} = 12.2\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{27.0\ \cancel{\text{g}}} = 0.452\ \text{mol Al}$$

d) 33.5 g of iron atoms

Answer:

$$\text{moles} = 33.5\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{55.8\ \cancel{\text{g}}} = 0.600\ \text{mol Fe}$$

e) 4.40 g of CO_2

Answer:

$$\text{moles} = 4.40\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{44.0\ \cancel{\text{g}}} = 0.100\ \text{mol}\ CO_2$$

f) 3.612×10^{24} atoms of oxygen

Answer:

$$\text{moles} = 3.612 \times 10^{24}\ \cancel{\text{atoms}} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \cancel{\text{atoms}}} = 6.00\ \text{mol O}$$

g) 1.20×10^{23} molecules of H_2O

Answer:

$$\text{moles} = 1.20 \times 10^{23}\ \cancel{\text{molecules}} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} = 0.199\ \text{mol}\ H_2O$$

Learning Activity 3.12: Converting Moles to Mass

Calculate the mass of each of the following:

a) 1.25 moles of NaOH

Answer:

NaOH = 40.0 g/mol

$$\text{mass} = \left(1.25\ \cancel{\text{mol}}\right)\left(\frac{40.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 50.0\ \text{g}$$

b) 0.0500 moles of KCl

Answer:

KCl = 74.6 g/mol

$$\text{mass} = \left(0.0500\ \cancel{\text{mol}}\right)\left(\frac{74.6\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 3.73\ \text{g}$$

c) 0.450 moles of $Mg_3(PO_4)_2$

Answer:

$Mg_3(PO_4)_2$ = 262.9 g/mol

$$\text{mass} = \left(0.450\ \cancel{\text{mol}}\right)\left(\frac{262.9\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 118\ \text{g}$$

d) 2.50 x 10-2 moles of $NaNO_3$

Answer:

$NaNO_3$ = 85.0 g/mol

$$\text{mass} = \left(2.50 \times 10^{-2}\ \cancel{\text{mol}}\right)\left(\frac{85.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 2.12\ \text{g}$$

e) 4.75 x 10^9 molecules of water

Answer:

H_2O = 18.0 g/mol

$$\text{mass} = 4.75 \times 10^{9}\ \cancel{\text{molecules}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} \times \frac{18.0\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 1.42 \times 10^{-13}\ \text{g}\ H_2O$$

f) 4.515 x 10^{23} atoms of phosphorous

Answer:

P = 31.0 g/mol

$$\text{mass} = 4.515 \times 10^{23}\ \cancel{\text{atoms}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{atoms}}} \times \frac{31.0\ g}{1\ \cancel{\text{mol}}}$$

$$= 23.2\ \text{g P}$$

g) 1.50 x 10^{23} molecules of NaCl

Answer:

NaCl = 58.5 g/mol

$$\text{mass} = 1.50 \times 10^{23}\ \cancel{\text{molecules}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} \times \frac{58.5\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 14.6\ \text{g NaCl}$$

Learning Activity 3.13: Converting between Moles and Number of Particles

1. Find the number of particles in each of the following:
   a) 1.20×10^{-15} moles of zinc

      Answer:

$$\text{particles} = \left(1.20 \times 10^{-15}\ \cancel{\text{mol}}\right)\left(\frac{6.02 \times 10^{23}\ \text{particles}}{1\ \cancel{\text{mol}}}\right)$$

$$= 7.22 \times 10^{8}\ \text{particles Zn}$$

   b) 4.50×10^{-7} moles of tin atoms

      Answer:

$$\text{particles} = \left(4.50 \times 10^{-7}\ \cancel{\text{mol}}\right)\left(\frac{6.02 \times 10^{23}\ \text{particles}}{1\ \cancel{\text{mol}}}\right)$$

$$= 2.71 \times 10^{17}\ \text{particles Sn}$$

   c) 5.75 moles of lead

      Answer:

$$\text{particles} = \left(5.75\ \cancel{\text{mol}}\right)\left(\frac{6.02 \times 10^{23}\ \text{particles}}{1\ \cancel{\text{mol}}}\right)$$

$$= 3.46 \times 10^{24}\ \text{particles Pb}$$

   d) 1.41 moles of water molecules

      Answer:

$$\text{particles} = \left(1.41\ \cancel{\text{mol}}\right)\left(\frac{6.02 \times 10^{23}\ \text{particles}}{1\ \cancel{\text{mol}}}\right)$$

$$= 8.49 \times 10^{23}\ \text{particles } H_2O$$

2. 75.6 g of ammonium sulfate will contain
   a) how many moles of ammonium sulfate?
      Answer:
      $(NH_4)_2SO_4$ = 132.1 g/mol

$$\text{moles} = 75.6\ \cancel{g} \times \frac{1\ \text{mol}}{132.1\ \cancel{g}} = 0.572\ \text{mol}\ (NH_4)_2\,SO_4$$

   b) how many moles of nitrogen?
      Answer:

$$\text{moles} = 0.572\ \text{mols}\ \cancel{(NH_4)_2\,SO_4} \times \frac{2\ N}{1\ \cancel{(NH_4)_2\,SO_4}} = 1.14\ \text{mol N}$$

3. Determine the number of moles of *each element* in
   a) 2.5 moles of water
      Answer:

$$\text{moles H} = 2.5\ \text{mols}\ H_2O \times \frac{2\ H}{1H_2O} = 5.0\ \text{mol H}$$

$$\text{moles O} = 2.5\ \text{moles}\ \cancel{H_2O} \times \frac{1\ O}{1\ \cancel{H_2O}} = 2.5\ \text{moles O}$$

   b) 5.00 moles of NH_4NO_3
      Answer:

$$\text{moles N} = 5.00\ \text{mols}\ \cancel{NH_4NO_3} \times \frac{2\ N}{1\ \cancel{NH_4NO_3}} = 10.0\ \text{mol N}$$

$$\text{moles H} = 5.00\ \text{mols}\ \cancel{NH_4NO_3} \times \frac{4\ H}{1\ \cancel{NH_4NO_3}} = 20.0\ \text{mol H}$$

$$\text{moles O} = 5.00\ \text{mols}\ \cancel{NH_4NO_3} \times \frac{3\ O}{1\ \cancel{NH_4NO_3}} = 15.0\ \text{mol O}$$

c) 10.0 g of $CaCO_3$

Answer:

$CaCO_3$ = 100.1 g/mol

$$\text{mol Ca} = 10.0\ \cancel{\text{g}}\ \cancel{\text{CaCO}_3} \times \frac{1\ \text{mol}}{100.1\ \cancel{\text{g}}} \times \frac{1\ \text{Ca}}{1\ \cancel{\text{CaCO}_3}} = 0.0999\ \text{mol Ca}$$

$$= 9.99 \times 10^{-2}\ \text{mol Ca}$$

$$\text{mol C} = 10.0\ \cancel{\text{g}}\ \cancel{\text{CaCO}_3} \times \frac{1\ \text{mol}}{100.1\ \cancel{\text{g}}} \times \frac{1\ \text{C}}{1\ \cancel{\text{CaCO}_3}} = 0.0999\ \text{mol C}$$

$$= 9.99 \times 10^{-2}\ \text{mol C}$$

$$\text{mol O} = 10.0\ \cancel{\text{g}}\ \cancel{\text{CaCO}_3} \times \frac{1\ \text{mol}}{100.1\ \cancel{\text{g}}} \times \frac{3\ \text{O}}{1\ \cancel{\text{CaCO}_3}} = 0.300\ \text{mol O}$$

d) 50.0 g of $Ba(OH)_2$

Answer:

$Ba(OH)_2$ = 171.3 g/mol

$$\text{mol Ba} = 50.0\ \cancel{\text{g}}\ \cancel{\text{Ba(OH)}_2} \times \frac{1\ \text{mol}}{171.3\ \cancel{\text{g}}} \times \frac{1\ \text{Ba}}{1\ \cancel{\text{Ba(OH)}_2}} = 0.292\ \text{mol Ba}$$

$$\text{mol O} = 50.0\ \cancel{\text{g}}\ \cancel{\text{Ba(OH)}_2} \times \frac{1\ \text{mol}}{171.3\ \cancel{\text{g}}} \times \frac{2\ \text{O}}{1\ \cancel{\text{Ba(OH)}_2}} = 0.584\ \text{mol O}$$

$$\text{mol H} = 50.0\ \cancel{\text{g}}\ \cancel{\text{Ba(OH)}_2} \times \frac{1\ \text{mol}}{171.3\ \cancel{\text{g}}} \times \frac{2\ \text{H}}{1\text{Ba(OH)}_2} = 0.584\ \text{mol H}$$

Learning Activity 3.14: Percent Composition, Empirical Formula, and Molecular Formula

1. Determine the percent composition of each element in the following.
   a) H_2SO_4

      Answer:

      H_2SO_4 = 98.1 g/mol

      $$\%H = \frac{(2 \times 1.0\ \cancel{g})}{98.1\ \cancel{g}} \times 100\% = 2.0\%$$

      $$\%S = \frac{(1 \times 32.1\ \cancel{g})}{98.1\ \cancel{g}} \times 100\% = 32.7\%$$

      $$\%O = \frac{(4 \times 16.0\ \cancel{g})}{98.1\ \cancel{g}} \times 100\% = 65.2\%$$

   b) $Fe(C_2H_3O_2)_3$

      Answer:

      $Fe(C_2H_3O_2)_3$ = 232.8 g/mol

      $$\%Fe = \frac{(1 \times 55.8\ \cancel{g})}{232.8\ \cancel{g}} \times 100\% = 24.0\%$$

      $$\%C = \frac{(6 \times 12.0\ \cancel{g})}{232.8\ \cancel{g}} \times 100\% = 30.9\%$$

      $$\%H = \frac{(9 \times 1.0\ \cancel{g})}{232.8\ \cancel{g}} \times 100\% = 3.9\%$$

      $$\%O = \frac{(6 \times 16.0\ \cancel{g})}{232.8\ \cancel{g}} \times 100\% = 41.2\%$$

c) $C_{12}H_{22}O_{11}$

Answer:

$C_{12}H_{22}O_{11}$ = 342.0 g/mol

$$\%C = \frac{(12 \times 12.0\ \cancel{g})}{342.0\ \cancel{g}} \times 100\% = 42.1\%$$

$$\%H = \frac{(22 \times 1.0\ \cancel{g})}{342.0\ \cancel{g}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11 \times 16.0\ \cancel{g})}{342.0\ \cancel{g}} \times 100\% = 51.5\%$$

2. What are the empirical formulas of the following compounds?

a) $C_2H_4(OH)_2$

Answer:

CH_3O

b) $H_2S_20_8$

Answer:

HSO_4

c) C_4H_{10}

Answer:

C_2H_5

d) C_2H_5OH

Answer:

C_2H_6O

3. Determine the empirical formula for each of the following.

a) 0.104 mol potassium, 0.052 mol carbon, and 0.156 mol oxygen

Answer:

$$K = \frac{0.104 \cancel{\text{mol}}}{0.052 \cancel{\text{mol}}} = 2$$

$$C = \frac{0.052 \cancel{\text{mol}}}{0.052 \cancel{\text{mol}}} = 1$$

$$O = \frac{0.156 \text{ mol}}{0.052 \cancel{\text{mol}}} = 3$$

K_2CO_3

b) 35.98% aluminum and 64.02% sulfur

Answer:

Assume 100 g

$$\text{moles Al} = 35.98 \cancel{\text{g}} \times \frac{1 \text{ mol}}{27.0 \cancel{\text{g}}} = 1.33 \text{ mol Al}$$

$$\text{moles S} = 64.02 \cancel{\text{g}} \times \frac{1 \text{ mol}}{32.1 \cancel{\text{g}}} = 1.99 \text{ mol S}$$

$$Al = \frac{1.33 \cancel{\text{mol}}}{1.33 \cancel{\text{mol}}} = 1$$

$$S = \frac{1.99 \cancel{\text{mol}}}{1.33 \cancel{\text{mol}}} = 1.5$$

$2(AlS_{1.5}) \rightarrow Al_2S_3$

c) 22.9% sodium, 21.5% boron, and 55.7% oxygen

Answer:

Assume 100 g

$$\text{moles Na} = 22.9\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{23.0\ \cancel{\text{g}}} = 0.996\ \text{mol Na}$$

$$\text{moles B} = 21.5\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{10.8\ \cancel{\text{g}}} = 1.99\ \text{mol B}$$

$$\text{moles O} = 55.7\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{16.0\ \cancel{\text{g}}} = 3.48\ \text{mol O}$$

$$\text{Na} = \frac{0.996\ \cancel{\text{mol}}}{0.996\ \cancel{\text{mol}}} = 1$$

$$\text{B} = \frac{1.99\ \cancel{\text{mol}}}{0.996\ \cancel{\text{mol}}} = 2$$

$$\text{O} = \frac{3.48\ \cancel{\text{mol}}}{0.996\ \cancel{\text{mol}}} = 3.5$$

$2(NaB_2O_{3.5}) \rightarrow Na_2B_4O_7$

d) 21.7% carbon, 9.6% oxygen, and 68.7% fluorine

Answer:

Assume 100 g

$$\text{moles C} = 21.7\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{12.0\ \cancel{\text{g}}} = 1.81\ \text{mol C}$$

$$\text{moles O} = 9.6\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{16.0\ \cancel{\text{g}}} = 0.60\ \text{mol O}$$

$$\text{moles F} = 68.7\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{19.0\ \cancel{\text{g}}} = 3.62\ \text{mol F}$$

$$\text{C} = \frac{1.81\ \text{mol}}{0.60\ \cancel{\text{mol}}} = 3$$

$$\text{O} = \frac{0.60\ \cancel{\text{mol}}}{0.60\ \cancel{\text{mol}}} = 1$$

$$\text{F} = \frac{3.62\ \cancel{\text{mol}}}{0.60\ \cancel{\text{mol}}} = 6$$

C_3OF_6

4. Given the empirical formula and molar mass, what is the molecular formula of each of the following compounds.
   a) HCO_2, molar mass = 90.0 g/mol
      Answer:
      HCO_2 = 45.0 g/mol
      $$x = \frac{90.0 \ \cancel{\text{g / mol}}}{45.0 \ \cancel{\text{g / mol}}} = 2$$
      Molecular formula = $H_2C_2O_4$

   b) C_2H_4O, molar mass = 88.0 g/mol
      Answer:
      C_2H_4O = 44.0 g/mol
      $$x = \frac{88.0 \ \cancel{\text{g / mol}}}{44.0 \ \cancel{\text{g / mol}}} = 2$$
      Molecular formula = $C_4H_8O_2$

   c) Na_2SiO_3, molar mass = 732.6 g/mol
      Answer:
      Na_2SiO_3 = 122.1 g/mol
      $$x = \frac{732.6 \ \cancel{\text{g / mol}}}{122.1 \ \cancel{\text{g / mol}}} = 6$$
      Molecular formula = $Na_{12}Si_6O_{18}$

5. Monosodium glutamate (MSG) is a flavour enhancer in certain foods that contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na. If MSG has a molar mass of 169.0 g/mol, what is its molecular formula?

Answer:

Assume 100 g

$$\text{moles C} = 35.51\ \cancel{g} \times \frac{1\ \text{mol}}{12.0\ \cancel{g}} = 2.959\ \text{mol}$$

$$\text{moles H} = 4.77\ \cancel{g} \times \frac{1\ \text{mol}}{1.0\ \cancel{g}} = 4.77\ \text{mol}$$

$$\text{moles O} = 37.85\ \cancel{g} \times \frac{1\ \text{mol}}{16.0\ \cancel{g}} = 2.366\ \text{mol}$$

$$\text{moles N} = 8.29\ \cancel{g} \times \frac{1\ \text{mol}}{14.0\ \cancel{g}} = 0.592\ \text{mol}$$

$$\text{moles Na} = 13.60\ \cancel{g} \times \frac{1\ \text{mol}}{23.0\ \cancel{g}} = 0.5913\ \text{mol}$$

$$C = \frac{2.959\ \cancel{\text{mol}}}{0.5913\ \cancel{\text{mol}}} = 5$$

$$H = \frac{4.77\ \cancel{\text{mol}}}{0.5913\ \cancel{\text{mol}}} = 8$$

$$O = \frac{2.366\ \cancel{\text{mol}}}{0.5913\ \cancel{\text{mol}}} = 4$$

$$N = \frac{0.592\ \cancel{\text{mol}}}{0.5913\ \cancel{\text{mol}}} = 1$$

$$Na = \frac{0.5913\ \cancel{\text{mol}}}{0.5913\ \cancel{\text{mol}}} = 1$$

The empirical formula of MSG is $C_5H_8O_4NNa$ = 169.0 g/mol.

$$x = \frac{169.0\ \cancel{\text{g / mol}}}{169.0\ \cancel{\text{g / mol}}} = 1$$

The molecular formula of MSG is $C_5H_8O_4NNa$.

6. Ibuprofen is a headache remedy with a molar mass about 206.0 g/mol. It contains 75.69% C, 8.80% H, and 15.51% O by mass. What is its molecular formula?

   Answer:

   Assume 100 g

$$\text{moles C} = 75.69\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{12.0\ \cancel{\text{g}}} = 6.308\ \text{mol}$$

$$\text{moles H} = 8.80\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{1.0\ \cancel{\text{g}}} = 8.80\ \text{mol}$$

$$\text{moles O} = 15.51\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{16.0\ \cancel{\text{g}}} = 0.9694\ \text{mol}$$

$$\text{C} = \frac{6.308\ \cancel{\text{mol}}}{0.9694\ \cancel{\text{mol}}} = 6.5$$

$$\text{H} = \frac{8.80\ \cancel{\text{mol}}}{0.9694\ \cancel{\text{mol}}} = 9$$

$$\text{O} = \frac{0.9694\ \cancel{\text{mol}}}{0.9694\ \cancel{\text{mol}}} = 1$$

   The empirical formula is $2(C_{6.5}H_9O) \rightarrow C_{13}H_{18}O_2$ = 206.0 g/mol.

$$x = \frac{206.0\ \cancel{\text{g/mol}}}{206.0\ \cancel{\text{g/mol}}} = 1$$

   Therefore, the molecular formula of Ibuprofen is $C_{13}H_{18}O_2$.

Grade 11 Chemistry (30S)

Midterm Practice Examination

Grade 11 Chemistry (30S)

Midterm Practice Examination

IInstructions

The midterm examination will be weighted as follows

Modules 1–3	100%

The format of the examination will be as follows:

Part A: Multiple Choice	49 x 1 = 49 marks
Part B: Fill-in-the-Blanks	15 x 1 = 15 marks
Part C: Short Answer	36 marks
Total Marks	100 marks

Include units with all answers as required.

Useful Information

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- some scrap paper
- a ruler
- a scientific or graphing calculator

You will have a maximum of 2.5 hours to complete your midterm exam.

Part A: Multiple Choice (49 Marks)

For each Multiple Choice question, shade in the circle that corresponds to your answer on the Bubble Sheet at the end of this exam. DO NOT circle your answers directly on the exam.

Physical Properties of Matter ***(20 marks)***

1. The pressure resulting from the collision of air molecules with objects is called
   a) Kinetic energy
   b) Atmospheric pressure
   c) Vapour pressure
   d) Sublimation

2. The pressure above a liquid in a sealed container caused by the collision of vaporized particles with the walls of their container is called
   a) Kinetic energy
   b) Vapour pressure
   c) Atmospheric pressure
   d) Sublimation

3. Which of the following statements does not agree with the kinetic theory of gases?
   a) Gas particles move in predictable patterns.
   b) Gas particles move independently of one another.
   c) Gas particles are spaced far apart from each other.
   d) Gas particles are in constant motion.

4. The average kinetic energy of the particles in a substance is
   a) Increased as the temperature of the substance increases.
   b) Unaffected by changes in the temperature of the substance.
   c) Increased as the temperature of the substance decreases.
   d) Equal to the total thermal energy absorbed by the substance.

5. What happens to the kinetic energy of the particles in a sample of gas as the temperature of the sample increases?
   a) It increases, then decreases
   b) It does not change
   c) It increases
   d) It decreases

6. Which of the following phase changes is NOT endothermic?
   a) Condensation
   b) Evaporation
   c) Melting
   d) Sublimation

7. The vaporization of a solid is also known as
   a) Condensation
   b) Deposition
   c) Evaporation
   d) Sublimation

8. The temperature at which the motion of particles theoretically ceases is
   a) 273 K
   b) –273 K
   c) 0 K
   d) 0°C

9. Which of the following is NOT a characteristic of liquids?
   a) Liquids have the ability to flow.
   b) The particles of a liquid are not attracted to each other.
   c) The particles of liquids are closer together than particles of gases.
   d) Liquids conform to the shape of their container.

10. Which are the first particles to evaporate from a liquid?
    a) Those with the lowest average kinetic energy.
    b) Those with the highest average kinetic energy.
    c) Those farthest from the surface of the liquid, regardless of kinetic energy.
    d) Those closest to the surface of the liquid, regardless of kinetic energy.

11. Which of these statements best explains why a liquid's rate of evaporation increases when the liquid is heated?
    a) The potential energy of the liquid increases, which in turn increases the rate of evaporation.
    b) The surface area of the liquid increases.
    c) More surface molecules have the energy required to overcome the attractive forces holding them in the liquid.
    d) The average kinetic energy of the liquid decreases.

12. If a liquid is sealed in an airtight container and kept at a constant temperature, how will its vapour pressure change over time?
    a) It rises at first, then falls over time.
    b) It rises at first, then remains constant.
    c) It remains constant.
    d) It rises continuously.

13. If a liquid is sealed in an airtight container, how will it be affected by an increase in temperature?
    a) The kinetic energy of the liquid particles will decrease.
    b) The atmospheric pressure above the liquid will increase.
    c) The vapour pressure of the liquid will increase.
    d) Fewer particles will escape the surface of the liquid.

14. In a dynamic equilibrium between the liquid state and the gaseous state, the rate of condensation
    a) Is greater than the rate of evaporation.
    b) Is less than the rate of evaporation.
    c) Is equal to the rate of evaporation.
    d) Has no effect on the rate of evaporation.

15. When gas particles escape from the surface of a liquid below the boiling point, this is called
    a) Evaporation
    b) Boiling
    c) Sublimation
    d) Condensation

16. When a liquid is at its normal boiling point, atmospheric pressure is
    a) 0 atm
    b) 0.5 atm
    c) 1 atm
    d) 2 atm

17. When the atmospheric pressure is equal to the vapour pressure of a liquid, the liquid will
    a) Condense
    b) Freeze
    c) Boil
    d) Melt

18. Which of the following characteristics is true of most solids?
    a) Solids are viscous.
    b) Solids are incompressible.
    c) Solids generally melt easily.
    d) Solids are made up of particles in rapid motion.

19. If you wanted to make water boil at 75ºC, instead of 100ºC, you would have to
    a) Add more heat.
    b) Increase air pressure.
    c) Decrease the volume of water you are boiling.
    d) Heat the water at a higher altitude.

20. Which of the following is not a phase change?
    a) Melting
    b) Evaporation
    c) Sublimation
    d) Diffusion

Gases and the Atmosphere ***(12 marks)***

21. A device used to measure atmospheric pressure is called a
    a) Manometer
    b) Barometer
    c) Thermometer
    d) Millibar

22. Gas pressure is usually measured with an instrument called a
    a) Manometer
    b) Barometer
    c) Pascal
    d) Thermometer

23. Which scientist discovered that the pressure of the atmosphere changed according to altitude?
    a) Huygens
    b) Avogadro
    c) Pascal
    d) Torricelli

24. It is possible for equal volumes of gases, at standard pressure and temperature, to contain equal numbers of particles because
    a) Gas particles are spaced far apart.
    b) Gas particles are large in size.
    c) The volume of a gas is inversely proportional to its mass.
    d) This is not possible.

25. Equal volumes of nitrogen and oxygen, at the same temperature and pressure, would
    a) Have the same mass.
    b) Contain a different number of particles.
    c) Contain the same number of particles.
    d) Have different average kinetic energies.

26. Which of the following best describes Boyle's Law?
    a) The volume of a gas is directly proportional to its temperature, if the pressure is kept constant.
    b) The volume of a gas varies inversely with pressure, at a constant temperature.
    c) The pressure of a gas is directly proportional to its temperature, if the volume is kept constant.
    d) At constant volume and temperature, the total pressure of a gas is equal to the sum of its partial pressures.

27. A weather balloon is heated from room temperature to 58ºC. As a result, the gas inside the weather balloon increases in volume. Which gas law explains this phenomenon?
    a) Gay-Lussac's Law
    b) Boyle's Law
    c) Charles' Law
    d) Combined Gas Law

28. Why does the pressure inside a container of gas increase if more gas is added to the container?
    a) There is an increase in the number of particles striking the wall of the container in the same period of time.
    b) An increase in gas causes an increase in temperature, which then increases pressure.
    c) As the volume of gas increases, the force of the collisions between particles and the container increases.
    d) As the gas pressure increases, the volume of gas decreases.

29. A gas is confined to a steel tank with a fixed volume. At 293 K, the gas exerts a pressure of 8.53 atm. After heating the tank, the pressure of the gas increases to 10.4 atm. What is the temperature of the heated gas?
    a) 357 K
    b) 326 K
    c) 240 K
    d) 926 K

30. According to Gay-Lussac's law
    a) Pressure is inversely proportional to temperature at a constant volume.
    b) Volume is inversely proportional to temperature at a constant pressure.
    c) Volume is directly proportional to temperature at a constant pressure.
    d) Pressure is directly proportional to temperature at a constant volume.

31. 4.0 L of a sample of gas at 1.0 atm of pressure is compressed into a 0.85 L tank. What is the pressure of the compressed gas, if the temperature remains constant?
    a) 0.15 atm
    b) 4.7 atm
    c) 0.21 atm
    d) 3.4 atm

32. A balloon is filled with 2.33 L of helium at 304 K. If the balloon is moved indoors where the temperature is 293 K, what will be the new volume of the balloon? Assume that pressure remains unchanged.
    a) 2.41 L
    b) 2.24 L
    c) 2.17 L
    d) 1.50 L

Chemical Reactions ***(17 marks)***

33. Calculate the number of moles in 21.2 g of hydrochloric acid (HCl).
    a) 0.581 moles
    b) 1.72 moles
    c) 21.0 moles
    d) 128 moles

34. What is the molar mass of $CaSO_4$?
    a) 108.1 g/mol
    b) 88.2 g/mol
    c) 136.2 g/mol
    d) 232.5 g/mol

35. How many moles are present in 165 g of manganese?
    a) 3.00
    b) 2.00
    c) 4.20
    d) 6.15

36. What is the formula mass of potassium chlorate, $KClO_3$?
    a) 122.6 μ
    b) 91.6 μ
    c) 193.6 μ
    d) 226.1 μ

37. How do the isotopes hydrogen–1 and hydrogen–2 differ?
    a) Hydogen–2 has one more electron than hydrogen–1.
    b) Hydrogen–2 has one more proton than hydrogen–1.
    c) Hydrogen–2 has one neutron, hydrogen–1 has none.
    d) Hydrogen–1 has one neutron, hydrogen–2 has 2 neutrons.

38. The fictional element Q has two naturally occurring isotopes with the following percent abundances: Q–20 is 25.0% abundant, and Q–22 is 75.0% abundant. What is the average atomic mass for Element Q?
    a) 20.5 g
    b) 21.0 g
    c) 21.5 g
    d) 42.0 g

39. What is the volume (in litres at STP) of 2.50 moles of carbon monoxide?
    a) 70.0 L
    b) 3.10 L
    c) 56.0 L
    d) 9.00 L

40. What is the number of moles in 500 L of He gas at STP?
    a) 0.05 moles
    b) 0.2 moles
    c) 20 moles
    d) 90 moles

41. What are the missing coefficients for the skeleton equation below?

$$Al_2(SO_4)_{3(aq)} + KOH_{(aq)} \rightarrow Al(OH)_{3(aq)} + K_2SO_{4(aq)}$$

    a) 1,6,2,3
    b) 2,3,1,1
    c) 1,3,2,3
    d) 4,6,3,2

42. Aluminum chloride and hydrogen gas are produced when strips of aluminum are placed in hydrochloric acid. What is the balanced equation for this reaction?
    a) $Al + 2HCl \rightarrow AlCl_2 + H_2$
    b) $Al + HCl_3 \rightarrow AlCl_3 + H$
    c) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$
    d) $H + AlCl \rightarrow Al + HCl$

43. What type of reaction was described in Question #42?
    a) Synthesis
    b) Single-replacement
    c) Double-replacement
    d) Decomposition

44. When the equation $Fe + Cl_2 \rightarrow FeCl_3$ is balanced, what is the coefficient for Cl_2?
    a) 4
    b) 3
    c) 2
    d) 1

45. If a synthesis reaction takes place between potassium and chlorine, what is the product?
   a) PCl_2
   b) PCl
   c) KCl_2
   d) KCl

46. What is the balanced equation for the decomposition of lead(IV)oxide?
   a) $2PbO \rightarrow 2Pb + O_2$
   b) $PbO \rightarrow Pb + O_2$
   c) $PbO_2 \rightarrow Pb + 2O_2$
   d) $PbO_2 \rightarrow Pb + O_2$

47. Which of the following is an empirical formula?
   a) Sb_2S_3
   b) $C_{12}H_{26}$
   c) $C_2H_8N_2$
   d) P_4O_{10}

48. The reaction type that involves two simple substances combining to produce a single more complex substance is known as
   a) Decomposition
   b) Single replacement
   c) Combustion
   d) Synthesis

49. You are given the following empirical formula: CH_2O. Which of the following may be the corresponding molecular formula?
   a) $C_6H_8O_6$
   b) $C_2H_{12}O_6$
   c) $C_6H_{12}O_6$
   d) $C_6H_{12}O_3$

Part B: Fill-in-the-Blanks (15 Marks)

Use the Word Bank at the end of this exam to help you complete the "Fill in the Blank" questions. As each blank is worth one mark, some questions will have a total value of two marks. Note that there are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use. The same term may be used more than once in this section.

Physical Properties of Matter *(5 marks)*

1. The temperature at which the vapour pressure of a liquid is equal to the external atmospheric pressure is called ____________________ point.
2. Collisions between gas molecules are assumed to be perfectly ____________________.
3. The ____________________ theory states that tiny particles in all forms of matter are in constant motion.
4. Unlike the other states of matter, ____________________ cannot flow.
5. The temperature at which the motion of particles theoretically ceases is known as ____________________ zero.

Gases and the Atmosphere *(5 marks)*

6. Methane is considered to be a ____________________ gas.
7. What we call ____________________ gas is actually not one gas, but a mixture of several naturally occurring gases in the atmosphere.
8. The ____________________ of natural gas has a direct effect on the carbon cycle, as the reaction produces both carbon dioxide and carbon monoxide.
9. Today, the air on Earth is mostly ____________________ and oxygen gases, but this was not always the case.
10. Blue-green algae produced oxygen gas through the process of ____________________.

Chemical Reactions *(5 marks)*

11. In a double-replacement reaction, the reactants are two ____________________ compounds.
12. One-twelfth the mass of one carbon atom equals one ____________________ mass unit.
13. The molar volume of a gas at ____________________ occupies 22.4 L.
14. In a ____________________ reaction, one of the reactants is oxygen gas.
15. Chemical equations must be balanced to satisfy the Law of ____________________ of mass.

Part C: Short Answer (37 Marks)

Answer each of the questions below using the space provided. Pay attention to the number of marks that each question is worth, as this may help you decide how much information to provide for full marks. For questions that involve calculations, show your work and check your final answer for the correct number of significant figures and the appropriate unit.

Physical Properties of Matter ***(5 marks)***

1. Describe the basic assumptions of the kinetic molecular theory of gases as far as:
   a) volume
   b) intermolecular forces
   c) collisions *(3 marks)*

2. Two jars of water are sealed and then stored at a temperature of 20ºC. One jar contains 75 mL of water, while the other contains 25 mL of water. Explain why, despite the difference in volume, the vapour pressure in both containers is the same. *(2 marks)*

Gases and the Atmosphere *(13 marks)*

3. A model hot air balloon has a volume of 35.0 L at a pressure of 100 kPa and a temperature of 80.0ºC. Calculate the new volume if the temperature is decreased to 15ºC (assume constant pressure). *(3 marks)*

4. A gas has a volume of 125 L at 325 kPa and 58.0°C. Use the Combined Gas Law to calculate the temperature in Celsius to produce a volume of 22.4 L at 101.3 kPa. Include the correct ratios for each part of the calculation, as well as a verbal prediction of each outcome. *(8 marks)*

5. Describe the strategies involved in one Air Quality Improvement initiative that you researched. *(2 marks, 1 mark per strategy discussed)*

Chemical Reactions *(18 marks)*

6. Name each of the following. *(4 marks)*

   a) Cu_3N

   b) $FeCO_3$

   c) $KMnO_4$

   d) P_3N_5

7. Write the chemical formula of the following compounds. *(5 marks)*

   a) mercury (II) iodide

   b) rubidium perchlorate

   c) calcium phosphate

   d) carbon tetrachloride

   e) bromine pentafluoride

8. List one isotope and identify how it is used. *(3 marks)*

9. What is the empirical formula of a compound that is 40.7% carbon, 54.2% oxygen, and 5.1% hydrogen? *(4 marks)*

10. Find the number of moles of calcium in a 0.400 g calcium supplement capsule. Show all of your calculations. *(2 marks)*

NOTES

Grade 11 Chemistry
Midterm Practice Examination

Word Bank

Use the following word bank to help you complete the "Fill-in-the-Blank" portion of your Midterm Examination. Note that there may be MORE terms here than you need, so read over the list carefully before choosing the terms that you want to use. You can also use certain words more than once.

absolute
ammonia
atomic
atoms
boiling
carbon dioxide
carbon monoxide
combustion
composition
conservation
decrease
elastic
equilibrium
evaporation
gas
gases
glucose
greenhouse
helium
increase
intermolecular
ionic
kinetic
liquids
mass
methane
mixture
mole
natural
neutron
neutrons
nitrogen
photosynthesis
plasma
processed
refined
replacement
solids
STP
sublimation
water vapour

Grade 11 Chemistry
Midterm Practice Examination
Bubble Sheet

Name: ______________________________________ ______________ / 49

For each Multiple Choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the exam.

	A	B	C	D		A	B	C	D		A	B	C	D		A	B	C	D
1.	❍	❍	❍	❍	14.	❍	❍	❍	❍	27.	❍	❍	❍	❍	40.	❍	❍	❍	❍
2.	❍	❍	❍	❍	15.	❍	❍	❍	❍	28.	❍	❍	❍	❍	41.	❍	❍	❍	❍
3.	❍	❍	❍	❍	16.	❍	❍	❍	❍	29.	❍	❍	❍	❍	42.	❍	❍	❍	❍
4.	❍	❍	❍	❍	17.	❍	❍	❍	❍	30.	❍	❍	❍	❍	43.	❍	❍	❍	❍
5.	❍	❍	❍	❍	18.	❍	❍	❍	❍	31.	❍	❍	❍	❍	44.	❍	❍	❍	❍
6.	❍	❍	❍	❍	19.	❍	❍	❍	❍	32.	❍	❍	❍	❍	45.	❍	❍	❍	❍
7.	❍	❍	❍	❍	20.	❍	❍	❍	❍	33.	❍	❍	❍	❍	46.	❍	❍	❍	❍
8.	❍	❍	❍	❍	21.	❍	❍	❍	❍	34.	❍	❍	❍	❍	47.	❍	❍	❍	❍
9.	❍	❍	❍	❍	22.	❍	❍	❍	❍	35.	❍	❍	❍	❍	48.	❍	❍	❍	❍
10.	❍	❍	❍	❍	23.	❍	❍	❍	❍	36.	❍	❍	❍	❍	49.	❍	❍	❍	❍
11.	❍	❍	❍	❍	24.	❍	❍	❍	❍	37.	❍	❍	❍	❍					
12.	❍	❍	❍	❍	25.	❍	❍	❍	❍	38.	❍	❍	❍	❍					
13.	❍	❍	❍	❍	26.	❍	❍	❍	❍	39.	❍	❍	❍	❍					

Periodic Table of the Elements

Atomic Number →	19
Symbol ←	**K**
Name →	Potassium
Relative Atomic Mass ←	39.1

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Ununhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Electronegativities Table

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 **Lanthanide Series**	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 **Actinide Series**	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Alphabetical Listing of the Elements and Their Atomic Masses

Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Actinium	(227)	Gold	197.0	Praseodymium	140.9
Aluminum	27.0	Hafnium	178.5	Promethium	(145)
Americium	(243)	Hassium	(265)	Protactinum	(231)
Antimony	121.7	Helium	4.0	Radium	(226)
Argon	39.9	Holmium	164.9	Radon	(222)
Arsenic	74.9	Hydrogen	1.0	Rhenium	186.2
Astatine	(210)	Indium	114.8	Rhodium	102.9
Barium	137.3	Iodine	126.9	Rubidium	85.5
Berkelium	(247)	Iridium	192.2	Ruthenium	101.1
Beryllium	9.0	Iron	55.8	Rutherfordium	(261)
Bismuth	209.0	Krypton	83.8	Samarium	150.4
Bohrium	(264)	Lanthanum	138.9	Scandium	45.0
Boron	10.8	Lawrencium	(257)	Seaborgium	(263)
Bromine	79.9	Lead	207.2	Selenium	79.0
Cadmium	112.4	Lithium	6.9	Silicon	28.1
Calcium	40.1	Lutetium	175.0	Silver	107.9
Californium	(251)	Magnesium	24.3	Sodium	23.0
Carbon	12.0	Manganese	54.9	Strontium	87.6
Cerium	140.1	Meitnerium	(266)	Sulfur	32.1
Cesium	132.9	Mendelevium	(256)	Tantalum	180.9
Chlorine	35.5	Mercury	200.6	Technetium	(98)
Chromium	52.0	Molybdenum	95.9	Tellurium	127.6
Cobalt	58.9	Neodymium	144.2	Terbium	158.9
Copernicium	(277)	Neon	20.2	Thallium	204.4
Copper	63.5	Neptunium	(237)	Thorium	232.0
Curium	(247)	Nickel	58.7	Thulium	168.9
Dubnium	(262)	Niobium	92.9	Tin	118.7
Dysprosium	162.5	Nitrogen	14.0	Titanium	47.9
Einsteinium	(254)	Nobelium	(259)	Tungsten	183.8
Erbium	167.3	Osmium	190.2	Uranium	238.0
Europium	152.0	Oxygen	16.0	Vanadium	50.9
Fermium	(257)	Palladium	106.4	Xenon	131.3
Fluorine	19.0	Phosphorus	31.0	Ytterbium	173.0
Francium	(223)	Platinum	195.1	Yttrium	88.9
Gadolinium	157.2	Plutonium	(244)	Zinc	65.4
Gallium	69.7	Polonium	(209)	Zirconium	91.2
Germanium	72.6	Potassium	39.1		

Names, Formulas, and Charges of Common Ions

Positive Ions (Cations)

Name	Symbol	Name	Symbol
aluminum	Al^{3+}	magnesium	Mg^{2+}
ammonium	NH_4^+	**manganese(II)**	**Mn^{2+}**
barium	Ba^{2+}	manganese(IV)	Mn^{4+}
cadmium	Cd^{2+}	mercury(I)	Hg_2^{2+}
calcium	Ca^{2+}	mercury(II)	Hg^{2+}
chromium(II)	Cr^{2+}	**nickel(II)**	**Ni^{2+}**
chromium(III)	**Cr^{3+}**	nickel(III)	Ni^{3+}
copper(I)	Cu^+	potassium	K^+
copper(II)	**Cu^{2+}**	silver	Ag^+
hydrogen	H^+	sodium	Na^+
iron(II)	Fe^{2+}	strontium	Sr^{2+}
iron(III)	**Fe^{3+}**	tin(II)	Sn^{2+}
lead(II)	**Pb^{2+}**	**tin(IV)**	**Sn^{4+}**
lead(IV)	Pb^{4+}	zinc	Zn^{2+}
lithium	Li^+		

continued

Negative Ions (Anions)

Name	Symbol	Name	Symbol
acetate	$C_2H_3O_2^-$ (CH_3COO^-)	nitrate	NO_3^-
azide	N_3^-	nitride	N^{3-}
bromide	Br^-	nitrite	NO_2^-
bromate	BrO_3^-	oxalate	$C_2O_4^{2-}$
carbonate	CO_3^{2-}	hydrogen oxalate	$HC_2O_4^-$
hydride	H^-	oxide	O^{2-}
hydrogen carbonate or bicarbonate	HCO_3^-	perchlorate	ClO_4^-
chlorate	ClO_3^-	permanganate	MnO_4^-
chloride	Cl^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}
chromate	CrO_4^{2-}	dihydrogen phosphate	$H_2PO_4^-$
citrate	$C_6H_5O_7^{3-}$	silicate	SiO_3^{2-}
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfate	HSO_4^-
fluoride	F^-	sulfide	S^{2-}
hydroxide	OH^-	hydrogen sulfide	HS^-
hypochlorite	ClO^-	sulfite	SO_3^{2-}
iodide	I^-	hydrogen sulfite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4^+ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

continued

Anions (Negative Ions)

1^{-} charge		1^{-} charge		2^{-} charge	
CH_3COO^{-} ($C_2H_3O_2^{-}$)	Acetate (or ethanoate)	HS^{-}	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^{-}	Bromate	OH^{-}	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^{-}	Bromide	IO_3^{-}	Iodate	O^{2-}	Oxide
ClO_3^{-}	Chlorate	I^{-}	Iodide	O_2^{2-}	Peroxide
Cl^{-}	Chloride	NO_3^{-}	Nitrate	SO_4^{2-}	Sulfate
ClO_2^{-}	Chlorite	NO_2^{-}	Nitrite	S^{2-}	Sulfide
CN^{-}	Cyanide	ClO_4^{-}	Perchlorate	SO_3^{2-}	Sulfite
F^{-}	Fluoride	IO_4^{-}	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^{-}	Hydride	MnO_4^{-}	Permanganate		
HCO_3^{-}	Hydrogen carbonate (or bicarbonate)	SCN^{-}	Thiocyanate	**3^{-} charge**	
				N^{3-}	Nitride
ClO^{-}	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^{-}	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Grade 11 Chemistry (30S)

Midterm Practice Examination

Answer Key

Grade 11 Chemistry (30S)

Midterm Practice Examination
Answer Key

IInstructions

The midterm examination will be weighted as follows

Modules 1–3	100%

The format of the examination will be as follows:

Part A: Multiple Choice	49 x 1 = 49 marks
Part B: Fill-in-the-Blanks	15 x 1 = 15 marks
Part C: Short Answer	36 marks
Total Marks	100 marks

Include units with all answers as required.

Useful Information

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- some scrap paper
- a ruler
- a scientific or graphing calculator

You will have a maximum of 2.5 hours to complete your midterm exam.

Part A: Multiple Choice (49 Marks)

For each Multiple Choice question, shade in the circle that corresponds to your answer on the Bubble Sheet at the end of this exam. DO NOT circle your answers directly on the exam.

Physical Properties of Matter *(20 marks)*

1. The pressure resulting from the collision of air molecules with objects is called
   a) Kinetic energy
   b) Atmospheric pressure
   c) Vapour pressure
   d) Sublimation

2. The pressure above a liquid in a sealed container caused by the collision of vaporized particles with the walls of their container is called
   a) Kinetic energy
   b) Vapour pressure
   c) Atmospheric pressure
   d) Sublimation

3. Which of the following statements does not agree with the kinetic theory of gases?
   a) Gas particles move in predictable patterns.
   b) Gas particles move independently of one another.
   c) Gas particles are spaced far apart from each other.
   d) Gas particles are in constant motion.

4. The average kinetic energy of the particles in a substance is
   a) Increased as the temperature of the substance increases.
   b) Unaffected by changes in the temperature of the substance.
   c) Increased as the temperature of the substance decreases.
   d) Equal to the total thermal energy absorbed by the substance.

5. What happens to the kinetic energy of the particles in a sample of gas as the temperature of the sample increases?
   a) It increases, then decreases
   b) It does not change
   c) It increases
   d) It decreases

6. Which of the following phase changes is NOT endothermic?
   a) Condensation
   b) Evaporation
   c) Melting
   d) Sublimation

7. The vaporization of a solid is also known as
   a) Condensation
   b) Deposition
   c) Evaporation
   d) Sublimation

8. The temperature at which the motion of particles theoretically ceases is
   a) 273 K
   b) –273 K
   c) 0 K
   d) 0°C

9. Which of the following is NOT a characteristic of liquids?
   a) Liquids have the ability to flow.
   b) The particles of a liquid are not attracted to each other.
   c) The particles of liquids are closer together than particles of gases.
   d) Liquids conform to the shape of their container.

10. Which are the first particles to evaporate from a liquid?
    a) Those with the lowest average kinetic energy.
    b) Those with the highest average kinetic energy.
    c) Those farthest from the surface of the liquid, regardless of kinetic energy.
    d) Those closest to the surface of the liquid, regardless of kinetic energy.

11. Which of these statements best explains why a liquid's rate of evaporation increases when the liquid is heated?
    a) The potential energy of the liquid increases, which in turn increases the rate of evaporation.
    b) The surface area of the liquid increases.
    c) More surface molecules have the energy required to overcome the attractive forces holding them in the liquid.
    d) The average kinetic energy of the liquid decreases.

12. If a liquid is sealed in an airtight container and kept at a constant temperature, how will its vapour pressure change over time?
    a) It rises at first, then falls over time.
    b) It rises at first, then remains constant.
    c) It remains constant.
    d) It rises continuously.

13. If a liquid is sealed in an airtight container, how will it be affected by an increase in temperature?
    a) The kinetic energy of the liquid particles will decrease.
    b) The atmospheric pressure above the liquid will increase.
    c) The vapour pressure of the liquid will increase.
    d) Fewer particles will escape the surface of the liquid.

14. In a dynamic equilibrium between the liquid state and the gaseous state, the rate of condensation
    a) Is greater than the rate of evaporation.
    b) Is less than the rate of evaporation.
    c) Is equal to the rate of evaporation.
    d) Has no effect on the rate of evaporation.

15. When gas particles escape from the surface of a liquid below the boiling point, this is called
    a) Evaporation
    b) Boiling
    c) Sublimation
    d) Condensation

16. When a liquid is at its normal boiling point, atmospheric pressure is
    a) 0 atm
    b) 0.5 atm
    c) 1 atm
    d) 2 atm

17. When the atmospheric pressure is equal to the vapour pressure of a liquid, the liquid will
    a) Condense
    b) Freeze
    c) Boil
    d) Melt

18. Which of the following characteristics is true of most solids?
    a) Solids are viscous.
    b) Solids are incompressible.
    c) Solids generally melt easily.
    d) Solids are made up of particles in rapid motion.

19. If you wanted to make water boil at 75ºC, instead of 100ºC, you would have to
    a) Add more heat.
    b) Increase air pressure.
    c) Decrease the volume of water you are boiling.
    d) Heat the water at a higher altitude.

20. Which of the following is not a phase change?
    a) Melting
    b) Evaporation
    c) Sublimation
    d) Diffusion

Gases and the Atmosphere ***(12 marks)***

21. A device used to measure atmospheric pressure is called a
    a) Manometer
    b) Barometer
    c) Thermometer
    d) Millibar

22. Gas pressure is usually measured with an instrument called a
    a) Manometer
    b) Barometer
    c) Pascal
    d) Thermometer

23. Which scientist discovered that the pressure of the atmosphere changed according to altitude?
    a) Huygens
    b) Avogadro
    c) Pascal
    d) Torricelli

24. It is possible for equal volumes of gases, at standard pressure and temperature, to contain equal numbers of particles because
    a) Gas particles are spaced far apart.
    b) Gas particles are large in size.
    c) The volume of a gas is inversely proportional to its mass.
    d) This is not possible.

25. Equal volumes of nitrogen and oxygen, at the same temperature and pressure, would
    a) Have the same mass.
    b) Contain a different number of particles.
    c) Contain the same number of particles.
    d) Have different average kinetic energies.

26. Which of the following best describes Boyle's Law?
    a) The volume of a gas is directly proportional to its temperature, if the pressure is kept constant.
    b) The volume of a gas varies inversely with pressure, at a constant temperature.
    c) The pressure of a gas is directly proportional to its temperature, if the volume is kept constant.
    d) At constant volume and temperature, the total pressure of a gas is equal to the sum of its partial pressures.

27. A weather balloon is heated from room temperature to 58ºC. As a result, the gas inside the weather balloon increases in volume. Which gas law explains this phenomenon?
    a) Gay-Lussac's Law
    b) Boyle's Law
    c) Charles' Law
    d) Combined Gas Law

28. Why does the pressure inside a container of gas increase if more gas is added to the container?
    a) There is an increase in the number of particles striking the wall of the container in the same period of time.
    b) An increase in gas causes an increase in temperature, which then increases pressure.
    c) As the volume of gas increases, the force of the collisions between particles and the container increases.
    d) As the gas pressure increases, the volume of gas decreases.

29. A gas is confined to a steel tank with a fixed volume. At 293 K, the gas exerts a pressure of 8.53 atm. After heating the tank, the pressure of the gas increases to 10.4 atm. What is the temperature of the heated gas?
    a) 357 K
    b) 326 K
    c) 240 K
    d) 926 K

30. According to Gay-Lussac's law
   a) Pressure is inversely proportional to temperature at a constant volume.
   b) Volume is inversely proportional to temperature at a constant pressure.
   c) Volume is directly proportional to temperature at a constant pressure.
   d) Pressure is directly proportional to temperature at a constant volume.

31. 4.0 L of a sample of gas at 1.0 atm of pressure is compressed into a 0.85 L tank. What is the pressure of the compressed gas, if the temperature remains constant?
   a) 0.15 atm
   b) 4.7 atm
   c) 0.21 atm
   d) 3.4 atm

32. A balloon is filled with 2.33 L of helium at 304 K. If the balloon is moved indoors where the temperature is 293 K, what will be the new volume of the balloon? Assume that pressure remains unchanged.
   a) 2.41 L
   b) 2.24 L
   c) 2.17 L
   d) 1.50 L

Chemical Reactions ***(17 marks)***

33. Calculate the number of moles in 21.2 g of hydrochloric acid (HCl).
   a) 0.581 moles
   b) 1.72 moles
   c) 21.0 moles
   d) 128 moles

34. What is the molar mass of $CaSO_4$?
   a) 108.1 g/mol
   b) 88.2 g/mol
   c) 136.2 g/mol
   d) 232.5 g/mol

35. How many moles are present in 165 g of manganese?
   a) 3.00
   b) 2.00
   c) 4.20
   d) 6.15

36. What is the formula mass of potassium chlorate, $KClO_3$?
   a) 122.6 μ
   b) 91.6 μ
   c) 193.6 μ
   d) 226.1 μ

37. How do the isotopes hydrogen–1 and hydrogen–2 differ?
   a) Hydogen–2 has one more electron than hydrogen–1.
   b) Hydrogen–2 has one more proton than hydrogen–1.
   c) Hydrogen–2 has one neutron, hydrogen–1 has none.
   d) Hydrogen–1 has one neutron, hydrogen–2 has 2 neutrons.

38. The fictional element Q has two naturally occurring isotopes with the following percent abundances: Q–20 is 25.0% abundant, and Q–22 is 75.0% abundant. What is the average atomic mass for Element Q?
   a) 20.5 g
   b) 21.0 g
   c) 21.5 g
   d) 42.0 g

39. What is the volume (in litres at STP) of 2.50 moles of carbon monoxide?
   a) 70.0 L
   b) 3.10 L
   c) 56.0 L
   d) 9.00 L

40. What is the number of moles in 500 L of He gas at STP?
    a) 0.05 moles
    b) 0.2 moles
    c) 20 moles
    d) 90 moles

41. What are the missing coefficients for the skeleton equation below?

$$Al_2(SO_4)_{3(aq)} + KOH_{(aq)} \rightarrow Al(OH)_{3(aq)} + K_2SO_{4(aq)}$$

    a) 1,6,2,3
    b) 2,3,1,1
    c) 1,3,2,3
    d) 4,6,3,2

42. Aluminum chloride and hydrogen gas are produced when strips of aluminum are placed in hydrochloric acid. What is the balanced equation for this reaction?
    a) $Al + 2HCl \rightarrow AlCl_2 + H_2$
    b) $Al + HCl_3 \rightarrow AlCl_3 + H$
    c) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$
    d) $H + AlCl \rightarrow Al + HCl$

43. What type of reaction was described in Question #42?
    a) Synthesis
    b) Single-replacement
    c) Double-replacement
    d) Decomposition

44. When the equation $Fe + Cl_2 \rightarrow FeCl_3$ is balanced, what is the coefficient for Cl_2?
    a) 4
    b) 3
    c) 2
    d) 1

45. If a synthesis reaction takes place between potassium and chlorine, what is the product?

a) PCl_2

b) PCl

c) KCl_2

d) KCl

46. What is the balanced equation for the decomposition of lead(IV)oxide?

a) $2PbO \rightarrow 2Pb + O_2$

b) $PbO \rightarrow Pb + O_2$

c) $PbO_2 \rightarrow Pb + 2O_2$

d) $PbO_2 \rightarrow Pb + O_2$

47. Which of the following is an empirical formula?

a) Sb_2S_3

b) $C_{12}H_{26}$

c) $C_2H_8N_2$

d) P_4O_{10}

48. The reaction type that involves two simple substances combining to produce a single more complex substance is known as

a) Decomposition

b) Single replacement

c) Combustion

d) Synthesis

49. You are given the following empirical formula: CH_2O. Which of the following may be the corresponding molecular formula?

a) $C_6H_8O_6$

b) $C_2H_{12}O_6$

c) $C_6H_{12}O_6$

d) $C_6H_{12}O_3$

Part B: Fill-in-the-Blanks (15 Marks)

Use the Word Bank at the end of this exam to help you complete the "Fill in the Blank" questions. As each blank is worth one mark, some questions will have a total value of two marks. Note that there are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use. The same term may be used more than once in this section.

Physical Properties of Matter ***(5 marks)***

1. The temperature at which the vapour pressure of a liquid is equal to the external atmospheric pressure is called ____________________ point. ***Boiling***
2. Collisions between gas molecules are assumed to be perfectly ____________________. ***Elastic***
3. The ____________________ theory states that tiny particles in all forms of matter are in constant motion. ***Kinetic***
4. Unlike the other states of matter, ____________________ cannot flow. ***Solids***
5. The temperature at which the motion of particles theoretically ceases is known as ____________________ zero. ***Absolute***

Gases and the Atmosphere ***(5 marks)***

6. Methane is considered to be a ____________________ gas. ***Greenhouse***
7. What we call ____________________ gas is actually not one gas, but a mixture of several naturally occurring gases in the atmosphere. ***Natural***
8. The ____________________ of natural gas has a direct effect on the carbon cycle, as the reaction produces both carbon dioxide and carbon monoxide. ***Combustion***
9. Today, the air on Earth is mostly ____________________ and oxygen gases, but this was not always the case. ***Nitrogen***
10. Blue-green algae produced oxygen gas through the process of ____________________. ***Photosynthesis***

Chemical Reactions *(5 marks)*

11. In a double-replacement reaction, the reactants are two ____________________ compounds. ***Ionic***
12. One-twelfth the mass of one carbon atom equals one ____________________ mass unit. ***Atomic***
13. The molar volume of a gas at ____________________ occupies 22.4 L. ***STP***
14. In a ____________________ reaction, one of the reactants is oxygen gas. ***Combustion***
15. Chemical equations must be balanced to satisfy the Law of ____________________ of mass. ***Conservation***

Part C: Short Answer (37 Marks)

Answer each of the questions below using the space provided. Pay attention to the number of marks that each question is worth, as this may help you decide how much information to provide for full marks. For questions that involve calculations, show your work and check your final answer for the correct number of significant figures and the appropriate unit.

Physical Properties of Matter ***(5 marks)***

1. Describe the basic assumptions of the kinetic molecular theory of gases as far as:
   a) volume
   b) intermolecular forces
   c) collisions *(3 marks)*

   Answer:

   a) The volume of a gas particle is much less than the total volume of the gas. In other words, most of the volume of a gas is empty space. *(1 mark)*

   b) The particles of gases are so spread out that the intermolecular forces between them are negligible. In other words, gas particles do not really attract or repel each other. *(1 mark)*

   c) All collisions between gas particles are perfectly elastic. This means that there is no kinetic energy lost or gained during these collisions. *(1 mark)*

2. Two jars of water are sealed and then stored at a temperature of 20ºC. One jar contains 75 mL of water, while the other contains 25 mL of water. Explain why, despite the difference in volume, the vapour pressure in both containers is the same. *(2 marks)*

   Answer:

   Vapour pressure is not dependent on volume of liquid, only on the kinetic energy of the molecules in the vapour. *(1 mark)*

   Since both jars contain water and are at the same temperature, their kinetic energies will be the same. *(1 mark)*

Gases and the Atmosphere *(13 marks)*

3. A model hot air balloon has a volume of 35.0 L at a pressure of 100 kPa and a temperature of 80.0ºC. Calculate the new volume if the temperature is decreased to 15ºC (assume constant pressure). *(3 marks)*

   Answer:

   Convert temperature to Kelvin:

   80.0°C + 273 = 353 K; 15°C + 273 = 288 K *(1 mark)*

   $$\left(\frac{288\ \cancel{K}}{353\ \cancel{K}}\right) \times 35.0\ L = 28.6\ L$$ *(2 marks)*

4. A gas has a volume of 125 L at 325 kPa and 58.0°C. Use the Combined Gas Law to calculate the temperature in Celsius to produce a volume of 22.4 L at 101.3 kPa. Include the correct ratios for each part of the calculation, as well as a verbal prediction of each outcome. *(8 marks)*

   Answer:

   58.0°C + 273 = 331 K *(1 mark)*

   Volume is decreased, so temperature should decrease. *(1 mark)*

   This is the correct ratio: $\frac{22.4\ L}{125\ L}$ *(1 mark)*

   $$\text{New Temperature} = 331\ K \times \left(\frac{22.4\ \cancel{L}}{125\ \cancel{L}}\right) \times ?$$ *(1 mark)*

   Since pressure decreases, temperature should also decrease. *(1 mark)*

   This is the correct ratio: $\frac{101.3\ kPa}{325\ kPa}$ *(1 mark)*

   $$\text{New Temperature} = 331\ K \times \left(\frac{22.4\ \cancel{L}}{125\ \cancel{L}}\right) \times \left(\frac{101.3\ \cancel{kPa}}{325\ \cancel{kPa}}\right) = 18.5\ K$$ *(1 mark)*

   Convert Kelvin to Celsius: 18.5 K − 273 = −254°C. *(1 mark)*

5. Describe the strategies involved in one Air Quality Improvement initiative that you researched. *(2 marks, 1 mark per strategy discussed)*

Answer:

Answers will vary. (Module 2, Lesson 1)

Your answer should be based on the research you did for Assignment 2.1: Air Quality Improvement Research. To get full marks, you must describe **two** of the strategies that are part of the initiative.

Note: Read the question carefully. You are not agreeing with or critiquing the initiative. You are just describing it.

Chemical Reactions *(18 marks)*

6. Name each of the following. *(4 marks)*

   a) Cu_3N

   Answer:

   copper (I) nitride

   b) $FeCO_3$

   Answer:

   iron (II) carbonate

   c) $KMnO_4$

   Answer:

   potassium permanganate

   d) P_3N_5

   Answer:

   triphosphorous pentanitride

7. Write the chemical formula of the following compounds. *(5 marks)*

a) mercury (II) iodide

Answer:

HgI_2

b) rubidium perchlorate

Answer:

$RbClO_4$

c) calcium phosphate

Answer:

$Ca_3(PO_4)_2$

d) carbon tetrachloride

Answer:

CCl_4

e) bromine pentafluoride

Answer:

BrF_5

8. List one isotope and identify how it is used. *(3 marks)*

Answer:

Answers will vary. (Module 3, Lesson 1)

The applications of various isotopes can be found on page 9 of Module 3. For full marks in this question, you must include:

- The isotope (^{131}I, ^{198}Au, etc.)
- The field it is used in (Traces, Treatment, or Climatology/Geology)
- What it is used for (measuring iodine levels, etc.)

A sample answer could be:

- In medical treatments, ^{60}Co is used to produce the gamma rays used to destroy brain tumours.

9. What is the empirical formula of a compound that is 40.7% carbon, 54.2% oxygen, and 5.1% hydrogen? *(4 marks)*

Answer:

$$\text{moles C} = 40.7\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.0\ \cancel{\text{g C}}} = 3.39\ \text{moles C}$$

$$\text{moles H} = 5.1\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.0\ \cancel{\text{g H}}} = 5.1\ \text{moles H}$$

$$\text{moles O} = 54.2\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = 3.39\ \text{moles O}$$

(0.5 marks x 3 = 1.5 marks)

$$\text{mole ratio of C} = \frac{3.39\ \text{mol}}{3.39\ \text{mol}} = 1$$

$$\text{mole ratio of H} = \frac{5.1\ \text{mol}}{3.39\ \text{mol}} = 1.5$$

$$\text{mole ratio of O} = \frac{3.39\ \text{mol}}{3.39\ \text{mol}} = 1$$

(0.5 marks x 3 = 1.5 marks)

$C_1H_{1.5}O_1 \rightarrow C_2H_3O_2$ *(1 mark)*

10. Find the number of moles of calcium in a 0.400 g calcium supplement capsule. Show all of your calculations. *(2 marks)*

Answer:

$$0.400\ \cancel{\text{g Ca}} \times \frac{1\ \text{mole Ca}}{40.1\ \cancel{\text{g Ca}}} = 9.98 \times 10^{-3}\ \text{mol (}$$

Notes

Grade 11 Chemistry
Midterm Practice Examination

Word Bank

Use the following word bank to help you complete the "Fill-in-the-Blank" portion of your Midterm Examination. Note that there may be MORE terms here than you need, so read over the list carefully before choosing the terms that you want to use. You can also use certain words more than once.

absolute
ammonia
atomic
atoms
boiling
carbon dioxide
carbon monoxide
combustion
composition
conservation
decrease
elastic
equilibrium
evaporation
gas
gases
glucose
greenhouse
helium
increase
intermolecular
ionic
kinetic
liquids
mass
methane
mixture
mole
natural
neutron
neutrons
nitrogen
photosynthesis
plasma
processed
refined
replacement
solids
STP
sublimation
water vapour

Grade 11 Chemistry
Midterm Practice Examination

Bubble Sheet

Name: ______________________________ ____________ / 49

For each Multiple Choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the exam.

	A	B	C	D		A	B	C	D		A	B	C	D		A	B	C	D
1.	❍	❍	❍	❍	14.	❍	❍	❍	❍	27.	❍	❍	❍	❍	40.	❍	❍	❍	❍
2.	❍	❍	❍	❍	15.	❍	❍	❍	❍	28.	❍	❍	❍	❍	41.	❍	❍	❍	❍
3.	❍	❍	❍	❍	16.	❍	❍	❍	❍	29.	❍	❍	❍	❍	42.	❍	❍	❍	❍
4.	❍	❍	❍	❍	17.	❍	❍	❍	❍	30.	❍	❍	❍	❍	43.	❍	❍	❍	❍
5.	❍	❍	❍	❍	18.	❍	❍	❍	❍	31.	❍	❍	❍	❍	44.	❍	❍	❍	❍
6.	❍	❍	❍	❍	19.	❍	❍	❍	❍	32.	❍	❍	❍	❍	45.	❍	❍	❍	❍
7.	❍	❍	❍	❍	20.	❍	❍	❍	❍	33.	❍	❍	❍	❍	46.	❍	❍	❍	❍
8.	❍	❍	❍	❍	21.	❍	❍	❍	❍	34.	❍	❍	❍	❍	47.	❍	❍	❍	❍
9.	❍	❍	❍	❍	22.	❍	❍	❍	❍	35.	❍	❍	❍	❍	48.	❍	❍	❍	❍
10.	❍	❍	❍	❍	23.	❍	❍	❍	❍	36.	❍	❍	❍	❍	49.	❍	❍	❍	❍
11.	❍	❍	❍	❍	24.	❍	❍	❍	❍	37.	❍	❍	❍	❍					
12.	❍	❍	❍	❍	25.	❍	❍	❍	❍	38.	❍	❍	❍	❍					
13.	❍	❍	❍	❍	26.	❍	❍	❍	❍	39.	❍	❍	❍	❍					

Periodic Table of the Elements

Atomic Number → 19
K ← Symbol
Name → Potassium
39.1 ← Relative Atomic Mass

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Ununhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Electronegativities Table

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 **Lanthanide Series**	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 **Actinide Series**	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Alphabetical Listing of the Elements and Their Atomic Masses

Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Actinium	(227)	Gold	197.0	Praseodymium	140.9
Aluminum	27.0	Hafnium	178.5	Promethium	(145)
Americium	(243)	Hassium	(265)	Protactinum	(231)
Antimony	121.7	Helium	4.0	Radium	(226)
Argon	39.9	Holmium	164.9	Radon	(222)
Arsenic	74.9	Hydrogen	1.0	Rhenium	186.2
Astatine	(210)	Indium	114.8	Rhodium	102.9
Barium	137.3	Iodine	126.9	Rubidium	85.5
Berkelium	(247)	Iridium	192.2	Ruthenium	101.1
Beryllium	9.0	Iron	55.8	Rutherfordium	(261)
Bismuth	209.0	Krypton	83.8	Samarium	150.4
Bohrium	(264)	Lanthanum	138.9	Scandium	45.0
Boron	10.8	Lawrencium	(257)	Seaborgium	(263)
Bromine	79.9	Lead	207.2	Selenium	79.0
Cadmium	112.4	Lithium	6.9	Silicon	28.1
Calcium	40.1	Lutetium	175.0	Silver	107.9
Californium	(251)	Magnesium	24.3	Sodium	23.0
Carbon	12.0	Manganese	54.9	Strontium	87.6
Cerium	140.1	Meitnerium	(266)	Sulfur	32.1
Cesium	132.9	Mendelevium	(256)	Tantalum	180.9
Chlorine	35.5	Mercury	200.6	Technetium	(98)
Chromium	52.0	Molybdenum	95.9	Tellurium	127.6
Cobalt	58.9	Neodymium	144.2	Terbium	158.9
Copernicium	(277)	Neon	20.2	Thallium	204.4
Copper	63.5	Neptunium	(237)	Thorium	232.0
Curium	(247)	Nickel	58.7	Thulium	168.9
Dubnium	(262)	Niobium	92.9	Tin	118.7
Dysprosium	162.5	Nitrogen	14.0	Titanium	47.9
Einsteinium	(254)	Nobelium	(259)	Tungsten	183.8
Erbium	167.3	Osmium	190.2	Uranium	238.0
Europium	152.0	Oxygen	16.0	Vanadium	50.9
Fermium	(257)	Palladium	106.4	Xenon	131.3
Fluorine	19.0	Phosphorus	31.0	Ytterbium	173.0
Francium	(223)	Platinum	195.1	Yttrium	88.9
Gadolinium	157.2	Plutonium	(244)	Zinc	65.4
Gallium	69.7	Polonium	(209)	Zirconium	91.2
Germanium	72.6	Potassium	39.1		

Names, Formulas, and Charges of Common Ions

Positive Ions (Cations)

Name	Symbol	Name	Symbol
aluminum	Al^{3+}	magnesium	Mg^{2+}
ammonium	NH_4^+	**manganese(II)**	$\mathbf{Mn^{2+}}$
barium	Ba^{2+}	manganese(IV)	Mn^{4+}
cadmium	Cd^{2+}	mercury(I)	Hg_2^{2+}
calcium	Ca^{2+}	mercury(II)	Hg^{2+}
chromium(II)	Cr^{2+}	**nickel(II)**	$\mathbf{Ni^{2+}}$
chromium(III)	$\mathbf{Cr^{3+}}$	nickel(III)	Ni^{3+}
copper(I)	Cu^+	potassium	K^+
copper(II)	$\mathbf{Cu^{2+}}$	silver	Ag^+
hydrogen	H^+	sodium	Na^+
iron(II)	Fe^{2+}	strontium	Sr^{2+}
iron(III)	$\mathbf{Fe^{3+}}$	tin(II)	Sn^{2+}
lead(II)	$\mathbf{Pb^{2+}}$	**tin(IV)**	$\mathbf{Sn^{4+}}$
lead(IV)	Pb^{4+}	zinc	Zn^{2+}
lithium	Li^+		

continued

Negative Ions (Anions)

Name	Symbol	Name	Symbol
acetate	$C_2H_3O_2^-$(CH_3COO^-)	nitrate	NO_3^-
azide	N_3^-	nitride	N^{3-}
bromide	Br^-	nitrite	NO_2^-
bromate	BrO_3^-	oxalate	$C_2O_4^{2-}$
carbonate	CO_3^{2-}	hydrogen oxalate	$HC_2O_4^-$
hydride	H^-	oxide	O^{2-}
hydrogen carbonate or bicarbonate	HCO_3^-	perchlorate	ClO_4^-
chlorate	ClO_3^-	permanganate	MnO_4^-
chloride	Cl^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}
chromate	CrO_4^{2-}	dihydrogen phosphate	$H_2PO_4^-$
citrate	$C_6H_5O_7^{3-}$	silicate	SiO_3^{2-}
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfate	HSO_4^-
fluoride	F^-	sulfide	S^{2-}
hydroxide	OH^-	hydrogen sulfide	HS^-
hypochlorite	ClO^-	sulfite	SO_3^{2-}
iodide	I^-	hydrogen sulfite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Common Ions

Cations (Positive Ions)

1^+ charge		**2^+ charge**		**3^+ charge**	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4^+ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

continued

Anions (Negative Ions)

1⁻ charge		**1⁻ charge**		**2⁻ charge**	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^-	Bromate	OH^-	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^-	Bromide	IO_3^-	Iodate	O^{2-}	Oxide
ClO_3^-	Chlorate	I^-	Iodide	O_2^{2-}	Peroxide
Cl^-	Chloride	NO_3^-	Nitrate	SO_4^{2-}	Sulfate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulfide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulfite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocyanate	**3⁻ charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Grade 11 Chemistry (30S)

Module 4: Stoichiometry

Module 4: Stoichiometry

Introduction

In the last module, you learned how to write and balance chemical equations. Now you will use that knowledge to interpret chemical reactions and ultimately predict the amount of products that result from a reaction. One type of problem you will learn to solve involves the limiting factor, a reactant that limits the amount of product that can be formed. Think of making sandwiches with a limited amount of bread. You can only make four sandwiches with eight bread slices, regardless of the fact you might have 16 slices of cheese. You will end this module by exploring the importance of using balanced chemical equations in industry.

Assignments in Module 4

When you have completed the assignments for Module 4, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 4.1	Interpreting a Balanced Equation
2	Assignment 4.2	Using the Molar Ratio
3	Assignment 4.3	Converting between Volume and Mass
4	Assignment 4.4	Solving Limiting Reactant Problems
5	Assignment 4.5	Limiting Reactant Investigation
6	Assignment 4.6	Stoichiometry Applications

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

Lesson 1: Interpreting a Balanced Equation (1.5 hours)

Lesson Focus

SLO C11-3-12: Interpret a balanced equation in terms of moles, mass, and volumes of gases.

Coefficients

In previous lessons you have balanced chemical equations, some of which contained polyatomic ions (e.g., $NaNO_3$). Balanced equations contain important information about a chemical reaction, including coefficients, the numbers placed before each chemical compound. What do these coefficients actually represent? They can represent moles, molecules, or volumes (of gas). You can use the coefficients to develop ratios that will calculate how the reactants will combine, and predict the amount of products that result from the reaction.

Stoichiometry is the use of ratios to determine the quantities of reactants used and products produced in a chemical reaction. Stoichiometry comes from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In other words, stoichiometry means measuring the reactants and products of a reaction, which is what you are really doing when you interpret a balanced chemical equation.

Balanced Chemical Equations

The coefficients in a balanced chemical equation show us in what proportions the reactants combine and products are made. Here's an example of a balanced equation:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

In this case, the coefficients represent the number of molecules of each compound involved in the reaction. You can write the previous chemical reaction as a word equation:

2 molecules of hydrogen gas	+	1 molecule of oxygen gas	→	2 molecules of water vapour

Since working with individual molecules is not practical, the coefficients can also be interpreted as moles of reactants and products, since the overall ratio still remains constant:

2 moles of hydrogen gas	+	1 mole of oxygen gas	→	2 moles of water vapour

If you consider the coefficients to be the number of moles that combine or are produced, they can then be called **molar coefficients**. You can now convert from moles to mass:

2 moles (2.0 g/mol) = 4.0 g of hydrogen gas	+	1 mole (32.0 g/mol) of oxygen gas	→	2 moles (18.0 g/mol = 36.0 g of water vapour

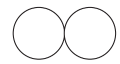
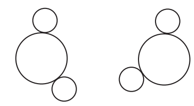
or moles to volume at STP (but only when reactants and products are gases!):

2 moles (22.4 L/mol) = 44.8 L of hydrogen gas	+	1 mole (22.4 L/mol) = 22.4 L of oxygen gas	→	2 moles (22.4 L/mol = 44.8 L of water vapour

Using the mole relationship, you can also relate moles to the number of atoms involved in the chemical equation.

2 moles (2 atoms/mol) = 4 atoms of hydrogen gas	+	1 mole (2 atoms/mol) = 2 atoms of oxygen gas	→	2 moles (2 atoms/mol) = 4 atoms of hydrogen gas
				2 moles (1 atom/mol) = 2 atoms of oxygen gas

All of the above scenarios outline interpretations of what balanced reaction coefficients can represent. It is important to note that mass (4.0 g + 32.0 g = 36.0 g) and atoms are conserved in every chemical reaction. However, in many reactions, molecules, formula units, moles, and volume are not conserved.

Stoichiometric Relationships in Chemical Reactions			
Interpretation of the Balanced Equation	$2H_{2(g)}$ +	$O_{2(g)}$ →	$2H_2O_{(g)}$
Representation of Reactants and Products			
Number of Molecules	2 molecules H_2	1 molecule of O_2	2 molecules of H_2O
Number of Atoms	4 atoms hydrogen	2 atoms oxygen	4 atoms hydrogen and 2 atoms oxygen (conserved)
Moles	2 moles of H_2	1 mole of O_2	2 moles of H_2O
Mass	4.0 g H_2	32.0 g O_2	36.0 g H_2O (conserved)
Volumes	44.8 L at STP	22.4 L at STP	44.8 L at STP

The Molar Ratio

A balanced chemical reaction is like a recipe. If you want to make 12 muffins, you can follow the original recipe. If you want to make 24 muffins, then you must double the recipe. Increasing the amount of reactants used will yield an increase in the amount of product formed. In order to increase the product formed and know how much reactant to use, it is necessary to determine ratios. Substances combine and are produced in distinct whole number ratios. You can determine these ratios using the coefficients from the balanced chemical equation.

Recall, in our example of the synthesis of water, that

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

What would you do in order to make 4 moles of water? If 2 moles of hydrogen combine with 1 mole of oxygen to produce 2 moles of water, then 4 moles (2 x 2) of hydrogen will combine with 2 moles (2 x 1) of oxygen to produce 4 moles (2 x 2) of water. The reaction would then be represented as:

$$4H_{2(g)} + 2O_{2(g)} \rightarrow 4H_2O_{(g)}$$

It's like doubling a cake recipe. Regardless of what amounts of reactants are present, they will always react in a 2 moles of H_2 for every 1 mole of O_2 ratio. We can represent this ratio in terms of a fraction:

$$\frac{2 \text{ moles } H_2}{1 \text{ mole } O_2}$$

The ratio of coefficients in a chemical equation is known as the **molar ratio** (it is also sometimes referred to as the **mole ratio**) for those two substances. In calculations, the molar ratio will help you convert between moles of reactants and moles of product, between moles of reactants, or between moles of products.

Learning Activity 4.1

Describing a Balanced Equation

1. How is a balanced equation similar to a recipe?
2. Chemical reactions can be described in terms of what quantities?
3. Interpret

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

   in terms of

   a) numbers of moles

   b) mass

   c) volumes of gases at STP

Lesson Summary

In this lesson, you learned how coefficients represent different quantities in chemical reactions. In many cases, the coefficients in a balanced chemical equation describe the number of moles of reactants and products involved in the reaction.

Next, you learned how to formulate the molar ratio. The molar ratio is the ratio of coefficients for two substances in a reaction and is used to determine the amounts of reactant needed or product formed in a chemical reaction. In the next lesson, you will solve stoichiometric problems involving moles, having been given the reactants and products in a balanced equation.

NOTES

Assignment 4.1

Interpreting a Balanced Equation *(9 marks)*

1. Interpret

   $2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

   in terms of

   a) numbers of molecules *(2 marks)*

   b) moles *(2 marks)*

   c) mass *(2 marks)*

2. The following equation shows the formation of aluminum oxide. From this equation, there are six different molar ratios that can be derived. Write any three of these molar ratios. *(3 marks)*

   $4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

NOTES

Lesson 2: Stoichiometry: Part 1 (2 hours)

Lesson Focus

SLO C11-3-13: Solve stoichiometry problems involving moles, mass, and volume, given the reactants and products in a balanced equation.
Include: heat of reaction problems

Lesson Introduction

In this lesson, you will learn how to use the molar ratio to calculate the amounts of reactants that combine and products that are formed in a chemical reaction. Using this important tool, you will solve stoichiometric problems involving conversions between moles and mass, based on a balanced chemical equation.

Solving Stoichiometry Problems

In the previous lesson, you learned that the coefficients in a balanced chemical equation can be used to create a ratio of moles that are being reacted or produced. It is therefore important that any quantity be converted to moles before using the molar ratio.

When solving stoichiometry problems, follow these steps:

Step 1: *If not already balanced, balance the chemical equation.*

Step 2: *If necessary, convert the given amount from mass or volume to moles.*

Step 3: *Use the molar ratio to calculate the moles of the required substance.*

Step 4: *If needed, convert the moles of the required substance to mass or volume.*

Mole-Mole Calculations

To solve stoichiometry problems, you must determine the amounts of reactants that are used and products that are formed in a chemical reaction. The molar ratio for the substances involved helps us determine quantities, in terms of moles. To make the molar ratio, always use:

$$\frac{\text{coefficient of substance needed}}{\text{coefficient of substance given}}$$

We can use a slightly altered set of the previous steps to solve stoichiometric problems involving moles:

Step 1: *Balance the chemical equation, if not already done.*

Step 2: *Determine the molar ratio for the substances involved.*

Step 3: *Multiply the given amount by the molar ratio.*

Example 1

For the following unbalanced equation,

$$H_{2(g)} + N_{2(g)} \rightarrow NH_{3(g)}$$

how many moles of ammonia will be produced from the reaction of 6.0 moles of hydrogen gas?

The overview of the process you will follow looks like this:
mol H_2 → mol NH_3

Step 1: *Balance the chemical equation, if not already done.*

$$3H_{2(g)} + 1N_{2(g)} \rightarrow 2NH_{3(g)}$$

Step 2: *Determine the molar ratio for the substances involved.*

The substance needed is ammonia and the substance given is hydrogen.

$$\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}$$

Note: When writing the molar ratio, include the number 1 if it is a coefficient in the balanced equation.

Step 3: *Multiply the given amount by the molar ratio.*

We are given 6.0 moles of H_2. Since the number of moles of hydrogen is given, hydrogen will be the denominator for the molar ratio.

$$\text{mol } NH_3 = 6.0 \ \cancel{\text{mol } H_2} \times \left(\frac{2 \text{ mol } NH_3}{3 \ \cancel{\text{mol } H_2}} \right) = 4.0 \text{ mol } NH_3$$

4.0 moles of ammonia will be produced from 6.0 moles of hydrogen gas.

Note that coefficients are not used in the calculation of significant digits.

Example 2

For the following balanced equation,

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

how many moles of water will be produced from 7.0 moles of oxygen gas?

Step 1: *Balance the chemical equation.*

This equation is already balanced.

Step 2: *Determine the molar ratio for the substances involved.*

The substance needed is water, and the substance given is oxygen. The molar ratio is therefore:

$$\frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2}$$

Step 3: *Multiply the given amount by the molar ratio.*

We are given 7.0 moles of O_2.

$$\text{mol } H_2O = 7.0 \ \cancel{\text{mol } O_2} \times \left(\frac{2 \text{ mol } H_2O}{1 \ \cancel{\text{mol } O_2}} \right) = 14 \text{ mol } H_2O$$

14 moles of water will be produced from 7.0 moles of oxygen.

Example 3

For the following unbalanced equation,

$$H_{2(g)} + N_{2(g)} \rightarrow NH_{3(g)}$$

how many moles of hydrogen gas will react with 0.350 moles of nitrogen gas?

Step 1: *Balance the chemical equation.*

$$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$$

Step 2: *Determine the molar ratio for the substances involved.*

The substance needed is hydrogen and the substance given is nitrogen. The molar ratio is

$$\frac{3 \text{ mol } H_2}{1 \text{ mol } N_2}$$

Step 3: *Multiply the given amount by the molar ratio.*

You are given 0.350 moles of N_2.

$$\text{mol } H_2 = 0.350 \cancel{\text{ mol } N_2} \times \left(\frac{3 \text{ mol } H_2}{1 \cancel{\text{ mol } N_2}}\right) = 1.05 \text{ mol } H_2$$

1.05 moles of hydrogen will react completely with 0.350 moles of nitrogen.

Example 4

In the following unbalanced equation, how many moles of aluminum are needed to form 3.70 moles of aluminum oxide, Al_2O_3?

$$Al_{(s)} + O_{2(g)} \rightarrow Al_2O_{3(s)}$$

Step 1: *Balance the chemical equation.*

$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

Step 2: *Determine the molar ratio for the substances involved.*

The substance needed is aluminum, and the substance given is aluminum oxide. The molar ratio is

$$\frac{4 \text{ mol Al}}{2 \text{ mol } Al_2O_3}$$

Step 3: *Multiply the given amount by the molar ratio.*

You are given 3.70 moles of Al_2O_3.

$$\text{mol Al} = 3.70\ \cancel{\text{mol Al}_2\text{O}_3} \times \left(\frac{4\ \text{mol Al}}{2\ \cancel{\text{mol Al}_2\text{O}_3}}\right) = 7.40\ \text{mol Al}$$

7.40 moles of aluminum are needed to form 3.70 moles of aluminum oxide.

Moles-Mass Problems

There are two types of problems that you will see involving moles and mass. Either

- the amount given is in moles and the amount requested is in grams, or
- the amount given is in grams and the amount needed is in moles

We can use the following set of steps to solve moles-mass stoichiometric problems:

Step 1: *Balance the equation, if not already done.*

Step 2: *If necessary, convert the given amount from mass to moles.*

Step 3: *Determine the molar ratio for the substances involved. Use the molar ratio to calculate the moles of the required substance.*

Step 4: *If needed, convert the moles of the required substance to mass.*

Example 1

What mass of NaCl is formed from the reaction between 3.20 moles of chlorine gas and an excess of sodium, according to the equation below?

$$Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$$

Excess sodium means there is more sodium than is needed to react with all the chlorine. In other words, there will sodium leftover at the end of this reaction.

The overview of the process you will follow looks like this:
mol Cl_2 → mol NaCl → g NaCl

Step 1: *Balance the equation, if not already done.*

$2Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$

Step 2: *If necessary, convert the given amount from mass to moles.*

As the number of moles of chlorine gas was already provided in the question, we can move directly to Step 3.

Step 3: *Determine the molar ratio for the substances involved. Use the molar ratio to calculate the moles of the required substance.*

Remember that the molar ratio is always the coefficient of the substance needed divided by the coefficient of the substance given. In this case, the moles of chlorine gas is given, so chlorine will be the denominator for the molar ratio.

$$\text{mol NaCl} = 3.20\ \cancel{\text{mol Cl}_2} \times \left(\frac{2\ \text{mol NaCl}}{1\ \cancel{\text{mol Cl}_2}}\right) = 6.40\ \text{mol NaCl}$$

Step 4: *Convert the moles of the required substance to mass.*

If you need to refresh your memory on mole calculations, take another look at Lessons 6, 7 and 8 from Module 3.

NaCl = 58.5 g/mol

$$\text{mass} = 6.40\ \cancel{\text{mol}} \times \left(\frac{58.5\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 374\ \text{g NaCl}$$

Alternatively, you can put Steps 3 and 4 together using the Factor-Label method:

$$\text{mass} = 3.20\ \cancel{\text{mol Cl}_2} \times \left(\frac{2\ \cancel{\text{mol}}\ \text{NaCl}}{1\ \cancel{\text{mol Cl}_2}}\right) \times \left(\frac{58.5\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 374\ \text{g NaCl}$$

374 g NaCl can be made from 3.20 moles of chlorine gas.

Example 2

Calculate the moles of hydrogen gas needed to make 15.0 g of ammonia.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Step 1: *Balance the chemical equation.*

This chemical equation is already balanced.

Step 2: *Convert the given amount from mass to moles.*

First, find the molar mass of ammonia. NH_3 = 17.0 g/mol

$$\text{mol} = 15.0\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{17.0\ \cancel{\text{g}}}\right) = 0.88235\ \text{mol}$$

Do not round at this point. Carry a few extra digits into the next calculation.

Step 3: *Determine the molar ratio for the substances involved. Use the molar ratio to calculate the moles of the required substance.*

Information about hydrogen is what you need to find, and therefore it will be the numerator of the molar ratio. Don't forget to round your final answer and use the correct number of significant digits.

$$\text{mol } H_2 = 0.88235\ \cancel{\text{mol } NH_3} \times \left(\frac{3\ \text{mol } H_2}{2\ \cancel{\text{mol } NH_3}}\right)$$
$$= 1.3235\ \text{mol } H_2 \approx 1.32\ \text{mol } H_2$$

or by using the Factor-Label method,

$$\text{mol } H_2 = 15.0\ \cancel{\text{g}}\ \cancel{NH_3} \times \left(\frac{1\ \cancel{\text{mol}}}{17.0\ \cancel{\text{g}}}\right) \times \left(\frac{3\ \text{mol } H_2}{2\ \cancel{\text{mol}}\ \cancel{NH_3}}\right)$$
$$= 1.3235\ \text{mol } H_2 \approx 1.32\ \text{mol } H_2$$

You need 1.32 moles of hydrogen to make 15.0 g of ammonia.

Mass-Mass Problems

It is difficult to directly measure the mass of a mole using a laboratory scale. As such, you will often need to convert the number of moles produced to a mass in grams. As usual, make sure that you start with a balanced chemical equation. The mass you determine can be that of either a product or a reactant. This type of problem solving uses the same set of four steps from the beginning of the lesson.

Example 1

Calculate the number of grams of NH_3 produced when 5.40 g of H_2 reacts with an excess of nitrogen. Here is the balanced chemical equation:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

The overview of the process you will follow looks like this: $g\ H_2 \rightarrow mol\ H_2 \rightarrow mol\ NH_3 \rightarrow g\ NH_3$

Step 1: *Balance the chemical equation.*

This chemical equation is already balanced.

Step 2: *Convert the given amount from mass to moles.*

First, find the molar mass of hydrogen: H_2 = 2.0 g/mol.

$$\text{mol} = 5.40\ \cancel{\text{g H}_2} \times \frac{1\ \text{mol H}_2}{2.0\ \cancel{\text{g H}_2}} = 2.70\ \text{mol H}_2$$

Step 3: *Determine the molar ratio for the substances involved. Use the molar ratio to calculate the moles of the required substance.*

Information about ammonia is what you need to find, and therefore ammonia will be the numerator of the molar ratio. Don't forget to round your final answer and practice using the correct number of significant digits.

$$\text{mol NH}_3 = 2.70\ \cancel{\text{mol H}_2} \times \left(\frac{2\ \text{mol NH}_3}{3\ \cancel{\text{mol H}_2}}\right) = 1.80\ \text{mol NH}_3$$

Step 4: *Convert the moles of the required substance to mass.*

The molar mass of NH_3 is 17.0 g/mol.

$$\text{mass} = 1.80\ \cancel{\text{mol NH}_3} \times \frac{17.0\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}} = 30.6\ \text{g NH}_3$$

As you have already seen in previous examples, all of these steps can be combined together using the Factor-Label method:

$$\begin{aligned} &\text{mass NH}_3 \\ &= 5.40\ \cancel{\text{g H}_2} \times \left(\frac{1\ \cancel{\text{mol H}_2}}{2.0\ \cancel{\text{g H}_2}}\right) \times \left(\frac{2\ \cancel{\text{mol NH}_3}}{3\ \cancel{\text{mol H}_2}}\right) \times \left(\frac{17.0\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}}\right) \\ &= 30.6\ \text{g NH}_3 \end{aligned}$$

Does the result make sense? There are three conversions you perform to get the final answer, so it may be difficult to estimate the outcome. However, since the molar mass of NH_3 is much greater than that of H_2, your answer should also have a greater mass than the given mass you started with. Here is another example to try.

Example 2

How many grams of acetylene (C_2H_2) are produced when 5.00 g of calcium carbide (CaC_2) is added to water? Here is the chemical equation:

$$CaC_{2(s)} + H_20_{(l)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(aq)}$$

Step 1: *Balance the chemical equation.*

$$CaC_{2(s)} + 2H_20_{(l)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(aq)}$$

Step 2: *Convert the given amount from mass to moles.*

The molar mass of calcium carbide, CaC_2, is 64.1 g/mol.

$$\text{mol} = 5.00\ \cancel{\text{g CaC}_2} \times \frac{1\ \text{mol CaC}_2}{64.1\ \cancel{\text{g CaC}_2}} = 0.0780\ \text{mol CaC}_2$$

Step 3: *Use the molar ratio to calculate the moles of the required substance.*

Information about acetylene is what you need to find, and therefore it will be the numerator of the molar ratio. Don't forget to round your final answer and use the correct number of significant digits.

$$\text{mol C}_2\text{H}_2 = 0.0780\ \cancel{\text{mol CaC}_2} \times \left(\frac{1\ \text{mol C}_2\text{H}_2}{1\ \cancel{\text{mol CaC}_2}}\right) = 0.0780\ \text{mol C}_2\text{H}_2$$

Step 4: *Convert the moles of the required substance to mass.*

The molar mass of acetylene is 26.0 g/mol.

$$\text{mass} = 0.0780\ \cancel{\text{mol } C_2H_2} \times \frac{26.0\ \text{g } C_2H_2}{1\ \cancel{\text{mol } C_2H_2}} = 2.03\ \text{g } C_2H_2$$

Learning Activity 4.2

Solving Stoichiometric Problems

Show your work in solving these stoichiometric problems, including the ratios you used. Try to determine the appropriate number of significant figures for your final answer. Don't forget to balance the equation first! Remember that molar masses have all been rounded and are NOT included in your significant figure determination.

Use the following unbalanced reaction to answer questions 1 to 3.

$$NH_3 + O_2 \rightarrow N_2 + H_2O$$

1. a) What are the molar coefficients of the balanced reaction?
   b) How many moles of oxygen gas will react exactly with 1.60 mol of ammonia?
   c) How many moles of each product will be generated by 1.60 mol of ammonia?
2. a) How many moles of oxygen gas will react with 0.750 mol of ammonia gas?
   b) How many moles of each product will be produced from the 0.750 mol of ammonia?
3. Determine the number of moles of water that would be produced from 2.50 mol of ammonia reacting with an excess of oxygen gas.
4. How many moles of H_2S can be burned by 0.750 moles of oxygen gas?

$$2H_2S + 3O_2 \rightarrow 2H_2O + 2SO_2$$

continued

Learning Activity 4.2 (continued)

5. How many moles of oxygen can be produced from 1.80 moles of $KClO_3$?

   $2KClO_3 \rightarrow 2KCl + 3O_2$

6. Calculate the mass of oxygen and hydrogen that will be formed by the decomposition of 4.50 g of water, according to the following reaction.

   $2H_2O \rightarrow 2H_2 + O_2$

   H_2O = 18.0 g/mol, H_2 = 2.0 g/mol, O_2 = 32.0 g/mol

7. Calculate the mass of aluminum oxide produced from 8.00 g of oxygen gas reacting with an excess of aluminum metal, according to the following reaction.

   $4Al + 3O_2 \rightarrow 2Al_2O_3$

   O_2 = 32.0 g/mol, Al_2O_3 = 102.0 g/mol

8. What mass of water is needed to react exactly with 2.30 g of NO_2 gas, and what mass of HNO_3 will be formed, according to the following reaction?

   $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

   NO_2 = 46.0 g/mol, H_2O = 18.0 g/mol, HNO_3 = 63.0 g/mol

9. What mass of tin (II) nitrate will be formed from the reaction of 25.2 g of nitric acid (HNO_3) and an excess of tin, according to the following reaction?

   $4Sn + 10HNO_3 \rightarrow 4Sn(NO_3)_2 + NH_4NO_3 + 3H_2O$

   HNO_3 = 63.0 g/mol, $Sn(NO_3)_2$ = 242.7 g/mol

10. What mass of HCl is required to form 14.2 g of Cl_2?

    $4HCl + O_2 \rightarrow 2H_2O + 2Cl_2$

    HCl = 36.5 g/mol; Cl_2 = 71.0 g/mol

continued

Learning Activity 4.2 (continued)

11. What mass of H_3PO_4 will react with 60.0 g of NaOH ?

    $H_3PO4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$

    $NaOH = 40.0$ g/mol, $H_3PO_4 = 98.0$ g/mol

12. How many moles of hydrogen gas are formed from 18.3 g of HCl ?

    $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

    $HCl = 36.5$ g/mol, $H_2 = 2.0$ g/mol

Lesson Summary

In this lesson, you used the following sequence of steps to solve problems involving conversions between moles and mass:

1. Balance the chemical equation, if it is not already done.
2. Convert the given quantity to moles.
3. Use the molar ratio to find the moles of the required substance.
4. Convert the moles of the required substance to mass.

In the next lesson, you will use a similar series of steps to solve problems involving conversions between mass and volume.

Assignment 4.2

Using the Molar Ratio *(27 marks)*

Solve the following stoichiometric problems, balancing the chemical equation when it is necessary. Show your work, including the molar ratio used in your calculation.

1. How many moles of oxygen are needed to burn 0.400 moles of C_8H_{18}? *(2 marks)*

   $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

2. How many moles of oxygen are needed to form 122 moles of Fe_2O_3? *(3 marks)*

   $Fe + O_2 \rightarrow Fe_2O_3$

3. How many moles of carbon dioxide are formed from 0.251 moles of CH_4? *(2 marks)*

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

continued

Assignment 4.2: Using the Molar Ratio (continued)

4. If 0.900 moles of CuO is reduced according to the equation: *(7 marks)*

   $CuO + NH_3 \rightarrow H_2O + N_2 + Cu$

   a) How many moles of water are formed?

   b) How many moles of N_2 are formed? What is the mass of the N_2 formed?

   $N_2 = 28.0$ g/mol

5. What mass of ammonia can be produced from 5.00 moles of H_2? *(3 marks)*

   $N_2 + 3H_2 \rightarrow 2NH_3$

   $NH_3 = 17.0$ g/mol

continued

Assignment 4.2: Using the Molar Ratio (continued)

6. How many moles of carbon dioxide are formed when 64.3 g of CH_4 burn? *(3 marks)*

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

   $CH_4 = 16.0$ g/mol

7. What mass of NO is formed when 3.00 moles of HNO_3 react with Cu? *(3 marks)*

   $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$

   $NO = 30.0$ g/mol

8. What mass of $KMnO_4$ is needed to produce 35.5 g of chlorine gas? *(4 marks)*

   $2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

   $Cl_2 = 71.0$ g/mol, $KMnO_4 = 158.0$ g/mol

NOTES

Lesson Focus

SLO C11-3-13: Solve stoichiometry problems involving moles, mass, and volume, given the reactants and products in a balanced equation.
Include: heat of reaction problems

Solving Stoichiometric Problems

In the previous lesson, you used a sequence of steps to convert between moles and mass. In this lesson, you will use a similar set of steps to solve stoichiometric problems converting between moles and volume:

1. Balance the chemical equation.
2. Convert the given quantity to moles.
3. Use the molar ratio to find the moles of the required substance.
4. Convert the moles of the required substance to mass or volume.

Finding the molar equation is an important step in solving stoichiometric problems. Remember, to find the molar ratio, always use

$$\frac{\text{coefficient of substance needed}}{\text{coefficient of substance given}}$$

Volume-Mass Problems

In these types of problems, quantities are given in terms of volume or mass, and the required amount is in terms of mass or volume.

Example 1

Given the following balanced equation, what volume of O_2 is needed to completely react with 25.0 g of FeS, at STP?

$$4FeS_{(s)} + 7O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 4SO_{2(g)}$$

The overview of the process you will follow looks like this:
g FeS → mol FeS → mol O_2 → L O_2

Step 1: *Balance the chemical equation.*

The equation is already balanced, so go to the next step.

Step 2: *Convert the given quantity to moles.*

To do this, we must find the molar mass of FeS and then convert the mass of FeS to moles of FeS. Do not round your answer at this point.

FeS = 87.9 g/mol

$$\text{mol} = 25.0 \cancel{\text{g}} \times \left(\frac{1 \text{ mol}}{87.9 \cancel{\text{g}}}\right) = 0.2844 \text{ mol FeS}$$

or

$$\text{mol} = \left(\frac{25.0 \cancel{\text{g}}}{87.9 \cancel{\text{g}}/\text{mol}}\right) = 0.2844 \text{ mol FeS}$$

Step 3: *Use the molar ratio to find the moles of the required substance.*

Since information about oxygen is what you need to find, oxygen will be the numerator of your molar ratio.

$$\text{mol } O_2 = 0.2844 \cancel{\text{mol FeS}} \times \left(\frac{7 \text{ mol } O_2}{4 \cancel{\text{mol FeS}}}\right) = 0.4977 \text{ mol } O_2$$

Step 4: *Convert the moles of the required substance to volume at STP.*

$$\text{vol } O_2 = 0.4977 \cancel{\text{mol}} \times \left(\frac{22.4 \text{ L}}{1 \cancel{\text{mol}}}\right) = 11.1 \text{ L } O_2$$

Or, if you like to combine the steps together:

$$\text{vol } O_2 = 25.0 \cancel{\text{g}} \ \cancel{\text{FeS}} \times \left(\frac{1 \ \cancel{\text{mol}}}{87.9 \ \cancel{\text{g}}}\right) \times \left(\frac{7 \ \cancel{\text{mol}} \ O_2}{4 \ \cancel{\text{mol}} \ \cancel{\text{FeS}}}\right) \times \left(\frac{22.4 \text{ L}}{1 \ \cancel{\text{mol}}}\right)$$

$$= 11.1 \text{ L } O_2$$

or

$$\text{vol } O_2 = \left(\frac{25.0 \ \cancel{\text{g}} \ \cancel{\text{FeS}}}{87.9 \ \cancel{\text{g}} / \cancel{\text{mol}}}\right) \times \left(\frac{7 \ \cancel{\text{mol}} \ O_2}{4 \ \cancel{\text{mol}} \ \cancel{\text{FeS}}}\right) \times \left(\frac{22.4 \text{ L}}{1 \ \cancel{\text{mol}}}\right) = 11.1 \text{ L } O_2$$

11.1 L of oxygen is needed to react with 25.0 g of FeS.

Example 2

According to the equation below, what mass of ammonia can be produced from 50.0 L of nitrogen gas, at STP?

$$N_{2(g)} + H_{2(g)} \rightarrow NH_{3(g)}$$

Step 1: *Balance the chemical equation.*

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Step 2: *Convert the given quantity to moles.*

Remember that 22.4 L is a constant, so it does not affect significant figure determination. Additionally, don't round until the last step.

$$\text{mol} = 50.0 \ \cancel{\text{L}} \times \left(\frac{1 \text{ mol}}{22.4 \ \cancel{\text{L}}}\right) = 2.232 \text{ mol}$$

or

$$\text{mol} = \left(\frac{50.0 \ \cancel{\text{L}}}{22.4 \ \cancel{\text{L}} / \text{mol}}\right) = 2.232 \text{ mol}$$

Step 3: *Use the molar ratio to find the moles of the required substance.*

Since information about ammonia is what you need to find, ammonia will be the numerator of your molar ratio.

$$\text{mol } NH_3 = 2.232 \ \cancel{\text{mol } N_2} \times \left(\frac{2 \text{ mol } NH_3}{1 \ \cancel{\text{mol } N_2}}\right) = 4.464 \text{ mol } NH_3$$

Step 4: *Convert the moles of the required substances to mass.*

The molar mass of ammonia is 17.0 g/mol.

$$\text{mass NH}_3 = 4.464\ \cancel{\text{mol}} \times \left(\frac{17.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 75.88\ \text{g} \approx 75.9\ \text{g NH}_3$$

Or, you can combine all of the steps together using the Factor-Label method:

$$\text{mass NH}_3 = 50.0\ \cancel{\text{L}}\ \cancel{\text{N}_2} \times \left(\frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}}\right) \times \left(\frac{2\ \text{NH}_3}{1\ \cancel{\text{N}_2}}\right) \times \left(\frac{17.0\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$= 75.88\ \text{g} \approx 75.9\ \text{g NH}_3$$

75.9 g of ammonia is produced from 50.0 L of nitrogen gas at STP.

Energy and Chemical Equations

Heat, light, and sound are all changes in energy that occur as a result of chemical reactions. Energy is either released (exothermic reaction) or absorbed (endothermic reaction). Earlier in this course, you learned that combustion is an example of an exothermic reaction, while melting is an example of an endothermic phase change.

Chemists show energy changes in chemical equations by including the energy term as a reactant or a product. For endothermic changes, the absorption of energy is shown by including the energy term as a reactant. An example of an endothermic change would be:

A + energy → B

For exothermic changes, the release of energy is represented by writing the energy term with the products. An example of an exothermic reaction would be:

A → B + energy

Energy and Stoichiometry

Now you know that energy terms can be included in chemical equations, and you know what an endothermic and an exothermic reaction might look like. You can also determine the amount of heat released or absorbed using stoichiometry. To calculate the amount of energy absorbed or released in a chemical reaction, you must treat the energy term the same as you do mass, moles, or volume. Likewise, the steps you followed to problem solve in previous sections will continue to be useful for this type of calculation.

Example 1

What quantity of heat is produced in the complete combustion of 60.2 g of ethane (C_2H_6), if the heat of combustion is 1560 kJ/mol of ethane?

$$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(g)}$$

The energy term 1560 kJ/mol can also be written as an **energy ratio**:

$$\frac{1560 \text{ kJ}}{1 \text{ mol}} \text{ or } \frac{1 \text{ mol}}{1560 \text{ kJ}}$$

The overview of the process you will follow looks like this: g C_2H_6 → mol C_2H_6 → kJ

Step 1: *Balance the chemical equation.*

The chemical equation is already balanced.

Step 2: *Convert the given quantity to moles.*

The molar mass of ethane is 30.0 g/mol.

$$\text{mol} = 60.2 \cancel{\text{g}} \times \left(\frac{1 \text{ mol}}{30.0 \cancel{\text{g}}}\right) = 2.007 \text{ mol}$$

or

$$\text{mol} = \left(\frac{60.2 \cancel{\text{g}}}{30.0 \cancel{\text{g}}/\text{mol}}\right) = 2.007 \text{ mol}$$

Step 3: *Use the energy ratio to determine the energy released.*

Note that 1560 kJ is the amount of energy released for every one mole of ethane. However, since you have 2.007 moles of ethane (the value you calculated in Step 2), you must multiply these two values together.

$$\text{energy} = 2.007\ \cancel{\text{mol}} \times \left(\frac{1560\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) = 3130\ \text{kJ}$$

Or, you can combine these two equations together and use the Factor-Label method.

$$\text{energy} = 60.2\ \cancel{\text{g}} \times \left(\frac{1\ \cancel{\text{mol}}}{30.0\ \cancel{\text{g}}}\right) \times \left(\frac{1560\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) = 3130\ \text{kJ}$$

Example 2

Given the equation

$$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O + 300.\ \text{kJ}$$

What mass of water is produced when 8000. kJ of energy is released by the reaction?

For every 6 moles of water, 300. kJ are released. We can represent this by the following energy ratios:

$$\frac{300\ \text{kJ}}{6\ \text{mol}\ H_2O} \quad \text{or} \quad \frac{6\ \text{mol}\ H_2O}{300\ \text{kJ}}$$

The overview of the process you will follow looks like this: kJ → moles H_2O → g H_2O

Step 1: *Balance the chemical equation.*

The equation is already balanced.

Step 2: *Use the energy ratio to determine the moles of the required substance.*

$$\text{mass}\ H_2O = 8000.\ \cancel{\text{kJ}} \times \left(\frac{6\ \text{mol}\ H_2O}{300.\ \cancel{\text{kJ}}}\right) = 160.\ \text{mol}\ H_2O$$

Step 3: *Convert the moles of the required substance to mass.*

The molar mass of water is 18.0 g/mol.

$$\text{mass } H_2O = (160.\ \cancel{\text{mol}}) \times \left(\frac{18.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 2880\ \text{g } H_2O$$

Or, you can combine all of the steps together using the Factor-Label method:

$$\text{mass } H_2O = 8000.\ \cancel{\text{kJ}} \times \left(\frac{6\ \cancel{\text{mol}}\ H_2O}{300.\ \cancel{\text{kJ}}}\right) \times \left(\frac{18.0\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 2880\ \text{g } H_2O$$

Example 3

129 kJ of energy is required to decompose 2 moles of $NaHCO_{3(s)}$. Calculate the amount of energy required to decompose 2.24 moles of $NaHCO_{3(s)}$, according to the following chemical equation:

$$2NaHCO_{3(s)} + 129\ \text{kJ} \rightarrow Na_2CO_{3(s)} + H_2O_{(g)} + CO_{2(g)}$$

Step 1: *Balance the chemical equation.*

The equation is already balanced.

Step 2: *Use the appropriate energy ratio to determine the required amount of energy.*

The energy ratios would be

$$\frac{129\ \text{kJ}}{2\ \text{mol NaHCO}_3} \quad \text{or} \quad \frac{2\ \text{mol NaHCO}_3}{129\ \text{kJ}}$$

$$2.24\ \cancel{\text{mol NaHCO}_3} \times \frac{129\ \text{kJ}}{2\ \cancel{\text{mol NaHCO}_3}} = 144\ \text{kJ}$$

Does this result make sense? If 129 kJ is required to decompose 2 mol $Na_2CO_{3(s)}$, then the decomposition of 2.24 mol should require about 10% more heat. The answer, 144 kJ, is consistent with this estimate and seems logical.

Learning Activity 4.3

Using the Molar Ratio to Calculate Volume

Show your work in solving these stoichiometric problems, including the ratios you used in your calculation. Balance the chemical equation when necessary and practice using the appropriate number of significant figures in your final answer.

1. What volume of each product could be made from 8.00 g of methane gas (CH_4) at STP, according to the following reaction?

   $CH_4 \rightarrow C_2H_2 + H_2$

2. What volume of CO_2 gas can be made from 11.2 L of CO gas and an excess of iron (III) oxide? Assume that temperature and pressure are kept constant at STP.

   $Fe_2O_3 + CO \rightarrow Fe + CO_2$

3. Given:

   $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$

   How many litres of nitrogen dioxide are produced when 34.0 L of oxygen gas reacts with an excess of nitrogen monoxide at STP?

4. $C_3H_8O_2 + 4O_2 \rightarrow 3CO_2 + 4H_2O$

   In the reaction above, the heat of combustion is 420.0 kJ/mole of $C_3H_8O_2$. When 125 L of $C_3H_8O_2$ react (at STP), how much energy would be released?

Lesson Summary

In this lesson, you learned how to solve problems where quantities are given in terms of volume or mass and the required amount is in terms of mass or volume. You also calculated the amount of heat released or absorbed in chemical reactions using stoichiometry. In doing so, you were able to identify a chemical reaction as either endothermic or exothermic. In the next lesson, you will learn about the limiting reactant in a chemical reaction. Your knowledge of stoichiometry will come in handy to help you calculate the mass of a product, given the reaction equation and reactant data.

Assignment 4.3

Converting between Volume and Mass *(13 marks)*

Show your work in solving these stoichiometric problems, including the ratios you used in your calculation. Don't forget to balance the chemical equation when necessary.

1. What volume of oxygen gas at STP can be made from 49.0 g of $KClO_3$ using the following reaction? *(5 marks)*

   $KClO_3$ = 122.6 g/mol

   $$KClO_3 \rightarrow KCl + O_2$$

2. What mass of glucose ($C_6H_{12}O_6$) would be required to make 5.60 L of CO_2 gas at STP, according to the following reaction? *(5 marks)*

   $C_6H_{12}O_6$ = 180.0 g/mol

   $$C_6H_{12}O_6 \rightarrow C_2H_5OH + CO_2$$

continued

Assignment 4.3: Converting between Volume and Mass (continued)

3. Calculate the energy released by the reaction when 1.00 g of H_2O_2 decomposes, according to the following equation. *(3 marks)*

   $2H_2O_2 \rightarrow 2H_2O + O_2 + 196.4\ kJ$

   $H_2O_2 = 34.0\ g/mol$

Lesson 4: Limiting Reactant (2 hours)

Lesson Focus

SLO C11-3-14: Identify the limiting reactant and calculate the mass of a product, given the reaction equation and reactant data.

Lesson Introduction

When determining the amount of product formed during a reaction, the amount of reactants that react according to the molar ratio is not always given. Sometimes there is some of one reactant leftover after the reaction is complete, which is referred to as an excess. In other cases, the reactants are used up completely in the chemical reaction. In this lesson, you will learn how to determine which reactant is leftover, and how the amount of product made by a reaction is affected by an insufficient quantity of any of the reactants.

Defining the Limiting Reactant

Imagine that you are having a group of your friends over to study chemistry, and you decide to make pizzas. You have enough cheese, sauce, and toppings for five pizzas, but you only have enough dough for two pizzas. How many pizzas can you make? Obviously, since you are limited by the amount of pizza dough you have, you can only make two pizzas, regardless of the quantity of toppings you have. The pizza dough is the limiting ingredient (or reactant) that determines the number of pizzas (or amount of product) that you can make.

In many reactions, one reactant becomes entirely consumed (or used up), and some of the other reactant is leftover. The reactant that becomes completely consumed is called the **limiting reactant**, also known as the **limiting reagent** or **limiting factor**. It is called the limiting reactant because when it is all gone no more product(s) can be created. This is why the amount of product is determined by the quantity of limiting reagent. In other words, the limiting reactant limits the amount of product that can be formed. The remaining, or leftover, reactants are called the **excess reactants**. Returning to our example above, the pizza dough is the limiting reactant, while the sauce and other toppings are the excess reactants.

You already know that every balanced chemical equation is essentially a chemical recipe. The coefficients you see in the balanced equation can be used to determine the *molar ratio*, which you have used in the last two lessons. The quantities in a balanced chemical equation can represent particles, volumes, mass or moles. Since you will need the molar ratio in your calculations, the first step in problem solving is to convert the amount of each reactant to moles. Here's another limiting reactant analogy using a balanced chemical equation:

Making a cheese sandwich (B_2C) requires 2 slices of bread (B) and 1 slice of cheese (C).

2 slices of bread + 1 slice of cheese $\rightarrow$ 1 cheese sandwich

$2B + 1C \rightarrow 1B_2C$

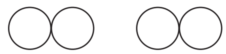
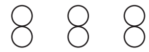
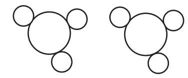
If you have 4 slices of bread and 2 slices of cheese, you can make exactly 2 sandwiches. Since there are no leftover ingredients, there is no limiting reactant. However, if you have 6 slices of bread and 2 slices of cheese, you can still only make 2 sandwiches. While there is enough bread to make 3 sandwiches, there is only enough cheese to make 2 sandwiches. The amount of cheese limits the number of sandwiches you can make. If this were a chemical reaction, the cheese would be the limiting reactant and the bread would be the excess reactant.

Example 1

Here's the balanced chemical equation for the preparation of ammonia, a common ingredient in fertilizers:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

In the reaction above, one mole of nitrogen reacts with three moles of hydrogen to produce two moles of ammonia. The recipe calls for 3 molecules of H_2 for every 1 molecule of N_2 used. As you can see from the table below, if you started the reaction with 2 moles of N_2 and 3 moles of H_2, you could still only make 2 moles of NH_3. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant.

Experimental Conditions			
	Reactants		Products
Before Reaction	2 molecules N_2 	3 molecules H_2 	0 molecules NH_3
After Reaction	1 molecule N_2 	0 molecules of H_2	2 molecules NH_3 

Remember that a coefficient can represent moles, molecules, volume, etc.

Solving Limiting Reactant Problems

There are several methods for solving limiting reactant problems. Here's one method that you can follow to solve these types of problems:

Step 1: *Balance the chemical equation, if it is not already balanced.*

Step 2: *If necessary, find the number of moles of each reactant.*

Step 3: *Using the molar coefficients, determine how much of one reactant is needed to use up the other.*

Step 4: I*f the amount from Step 3 is more than what is given of that reactant in Step 2, that reactant is the limiting reactant.*

Step 5: *Use the molar coefficients and the given amount of the limiting reactant to determine the moles of product.*

Step 6: *Convert moles of product to mass or volume.*

Step 7: *Determine the moles of excess reactant remaining after the reaction.*

Step 8: *Determine the mass of the excess reactant.*

Example 1

According to the equation below, how many moles of NaCl can be formed from a reaction between 6.70 moles of sodium and 3.20 moles of chlorine gas?

$$2Na(s) + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Notice that amounts are given for <u>both</u> reactants. This is an indication that the limiting factor must be determined.

Step 1: *Balance the chemical equation.*

The chemical equation is already balanced.

Step 2: *Find the number of moles of each reactant.*

This information was provided in the question.

Step 3: *Using the molar coefficients, determine how much of one reactant is needed to use up the other.*

You can start by choosing one of the reactants. It does not matter which reactant you begin with, as it will not change the final answer. Let's find how much chlorine gas is needed to use up the 6.70 moles of sodium.

$$\text{mol } Cl_2 = 6.70 \ \cancel{\text{mol Na}} \times \left(\frac{1 \text{ mol } Cl_2}{2 \ \cancel{\text{mol Na}}}\right) = 3.35 \text{ mol } Cl_2$$

Step 4: *If the amount from Step 3 is more than what is given of that reactant in Step 2, that reactant is the limiting reactant.*

According to the calculation involving the molar ratio, you need 3.35 moles of chlorine to use up all 6.70 moles of sodium. From the question, we are only given 3.20 moles of chlorine (3.20 – 3.35 = –0.15 moles). Since we do not have enough chlorine to use up the sodium, *chlorine gas is the limiting reactant* and sodium is therefore the excess reactant.

Note: Alternatively, you could begin with the sodium in Step 3:

$$\text{mol Na} = 3.20 \ \cancel{\text{mol } Cl_2} \times \left(\frac{2 \text{ mol Na}}{1 \ \cancel{\text{mol } Cl_2}}\right) = 6.40 \text{ mol Na}$$

According to the calculation, you would need 6.40 moles of sodium to use up all of the given chlorine. Since you are given 6.70 moles of sodium, there is more than you need (6.70 – 6.40 = +0.30 moles). This means that when 6.40 moles of sodium are used up, all of the chlorine will be gone, so *chlorine is still the limiting reactant*.

Step 5: *Use the molar coefficients and the given amount of the limiting reactant to determine the moles of product.*

In this step, you must use the given amount of chlorine, NOT the calculated amount in Step 3. The calculated amount is a hypothetical amount that you don't actually have.

$$\text{mol NaCl} = 3.20\ \cancel{\text{mol Cl}_2} \times \left(\frac{2\ \text{mol NaCl}}{1\ \cancel{\text{mol Cl}_2}}\right) = 6.40\ \text{mol NaCl}$$

6.40 moles of NaCl can be made from 6.70 moles of sodium and 3.20 moles of chlorine gas.

Example 2

How many grams of ammonia can be made from a reaction between 3.50 g of hydrogen gas and 18.0 g of nitrogen gas, according to the equation below? What mass of excess reactant remains?

$$N_{2(g)} + H_{2(g)} \rightarrow NH_{3(g)}$$

Again, quantities are given for both reactants, so this is a limiting reactant question.

Step 1: *Balance the chemical equation.*

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Step 2: *Find the number of moles of each reactant.*

The amounts are given in grams, which represents mass. You must first convert mass to moles:

$$\text{mol H}_2 = 3.50\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{2.0\ \cancel{\text{g}}}\right) = 1.75\ \text{mol H}_2$$

or

$$\text{mol H}_2 = \left(\frac{3.50\ \cancel{\text{g}}}{2.0\ \cancel{\text{g}}/\text{mol}}\right) = 1.75\ \text{mol H}_2$$

$$\text{mol N}_2 = 18.0\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{28.0\ \cancel{\text{g}}}\right) = 0.643\ \text{mol N}_2$$

or

$$\text{mol N}_2 = \left(\frac{18.0\ \cancel{\text{g}}}{28.0\ \cancel{\text{g}}/\text{mol}}\right) = 0.643\ \text{mol N}_2$$

Step 3: *Using the molar coefficients, determine how much of one reactant is needed to use up the other.*

You can choose any reactant to start with—say, hydrogen gas. Use the molar coefficients to calculate how many moles of nitrogen will use up all of the hydrogen.

$$\text{mol } N_2 = 1.75 \ \cancel{\text{mol } H_2} \times \left(\frac{1 \text{ mol } N_2}{3 \ \cancel{\text{mol } H_2}} \right) = 0.583 \text{ mol } N_2$$

Step 4: *If the amount from Step 3 is more than what is given of that reactant in Step 2, that reactant is the limiting reactant.*

You would need 0.583 moles of nitrogen to use up 1.75 moles of hydrogen. You are given 0.643 moles of nitrogen—more than enough. This means the hydrogen is used up first, so *hydrogen is the limiting reactant*.

You could also start with the nitrogen in Step 3:

$$\text{mol } H_2 = 0.643 \ \cancel{\text{mol } N_2} \times \left(\frac{3 \text{ mol } H_2}{1 \ \cancel{\text{mol } N_2}} \right) = 1.93 \text{ mol } H_2$$

Since there are only 1.75 moles of hydrogen, it is the limiting reactant, as you would need 1.93 moles of hydrogen to use up all of the nitrogen.

Step 5: *Use the molar coefficients and the given amount of the limiting reactant to determine the moles of product. Don't round off yet.*

$$\text{moles } NH_3 = 1.75 \ \cancel{\text{mol } H_2} \times \left(\frac{2 \text{ mol } NH_3}{3 \ \cancel{\text{mol } H_2}} \right) = 1.167 \text{ mol } NH_3$$

Step 6: *Convert moles of product to mass.*

NH_3 = 17.0 g/mol

$$\text{mass } NH_3 = 1.167 \ \cancel{\text{mol}} \times \frac{17.0 \text{ g}}{1 \ \cancel{\text{mol}}} = 19.8 \text{ g } NH_3$$

Step 7: *Determine the moles of excess reactant remaining after the reaction.*

Since hydrogen is the limiting reactant, nitrogen is the excess reactant. To find the moles of excess reactant remaining, you must calculate how much excess reactant is used up.

$$\text{mol } N_2 = 1.75\ \cancel{\text{mol } H_2} \times \left(\frac{1\text{ mol } N_2}{3\ \cancel{\text{mol } H_2}}\right) = 0.583\text{ mol } N_2$$

Moles of excess reactant remaining after the reaction

= initial moles of excess reactant – final moles of excess reactant

= 0.643 moles – 0.583 moles = 0.060 moles of N_2 in excess

Step 8: *Determine the mass of the excess reactant.*

N_2 = 28.0 g/mol

$$\text{mass} = 0.060\ \cancel{\text{mol}} \times \frac{28.0\text{ g}}{1\ \cancel{\text{mol}}} = 1.68\text{ g of } N_2 \text{ in excess}$$

19.8 g of ammonia is made from the reaction between 3.50 g of hydrogen and 18.0 g of nitrogen, with 1.68 g of nitrogen leftover after the reaction is complete.

Example 3

Propane burns according to the reaction below:

$$C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

What mass of carbon dioxide is formed by the reaction of 75.0 g of propane and 150.0 L of oxygen at STP?

As there are two amounts of reactants given (75.0 g of propane and 150.0 L of oxygen gas), we must first determine the limiting reactant.

Step 1: *Balance the chemical equation.*

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

Step 2: *Find the number of moles of each reactant.*

C_3H_8 = 44.0 g/mol

$$\text{mol } C_3H_8 = 75.0\ \cancel{\text{g}} \times \frac{1\text{ mol}}{44.0\ \cancel{\text{g}}} = 1.70\text{ mol } C_3H_8$$

or

$$\text{mol } C_3H_8 = \frac{75.0\ \cancel{\text{g}}}{44.0\ \cancel{\text{g}}/\text{mol}} = 1.70\text{ mol } C_3H_8$$

$$\text{mol } O_2 = 150.0 \cancel{\text{L}} \times \frac{1 \text{ mol}}{22.4 \cancel{\text{L}}} = 6.696 \text{ mol } O_2$$

or

$$\text{mol } O_2 = \frac{150.0 \cancel{\text{L}}}{22.4 \cancel{\text{L}}/\text{mol}} = 6.696 \text{ mol } O_2$$

Step 3: *Using the molar coefficients, determine how much of one reactant is needed to use up the other.*

You can use any reactant to start, say, oxygen gas. Therefore, you must find out how many moles of oxygen gas will react with all of the propane.

$$\text{mol } O_2 = 1.70 \cancel{\text{mol } C_3H_8} \times \frac{5 \text{ mol } O_2}{1 \cancel{\text{mol } C_3H_8}} = 8.50 \text{ mol } O_2$$

Step 4: *If the amount from Step 3 is more than what is given of that reactant in Step 2, that reactant is the limiting reactant.*

You are not given 8.50 moles of oxygen gas, so the oxygen is the limiting reactant.

Similarly, you could calculate the moles of propane that will react with all of the oxygen gas:

$$\text{mol } C_3H_8 = 6.696 \cancel{\text{mol } O_2} \times \frac{1 \text{ mol } C_3H_8}{5 \cancel{\text{mol } O_2}} = 1.339 \text{ mol } C_3H_8$$

You are given more propane than what is needed to consume all of the oxygen gas. The oxygen gas will be used up first, so oxygen is the limiting reactant.

Step 5: *Use the molar coefficients and the given amount of the limiting reactant to determine the moles of product.*

$$\text{mol } CO_2 = 6.696 \cancel{\text{mol } O_2} \times \frac{3 \text{ mol } CO_2}{5 \cancel{\text{mol } O_2}} = 4.018 \text{ mol } CO_2$$

Step 6: *Convert moles of product to mass.*

CO_2 = 44.0 g/mol

$$\text{mass} = 4.018 \cancel{\text{mol}} \times \frac{44.0 \text{ g}}{1 \cancel{\text{mol}}} = 176.8 \text{ g } CO_2$$

Steps 5 and 6 can be condensed together using the Factor-Label method:

$$\text{mass } CO_2 = 6.696 \ \cancel{\text{mol } O_2} \times \frac{3 \ \cancel{\text{mol}} \ CO_2}{5 \ \cancel{\text{mol } O_2}} \times \frac{44.0 \text{ g}}{1 \ \cancel{\text{mol}}} = 176.8 \text{ g } CO_2$$

In order to solidify your understanding of limiting versus excess reactants (and if you have access to the Internet), you can check out

- www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_03&folder=limiting_reactants
- the Limiting Reagent video http://cwx.prenhall.com/petrucci/medialib/media_ portfolio/04.html
- the Limiting Reagent flash animation www.mhhe.com/physsci/chemistry/essentialchemistry/flash/flash.mhtml.

Learning Activity 4.4

Limiting Factor and Excess Reactant

1. In each of the following questions, identify the *limiting factor* and the *excess reactant*. Use the following balanced equation:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

   a) 2 molecules of H_2 and 2 molecules of O_2
   b) 10 molecules of H_2 and 4 molecules of O_2
   c) 50 molecules of H_2 and 20 molecules of O_2
   d) 100 molecules of H_2 and 75 molecules of O_2
   e) 100 molecules of H_2 and 25 molecules of O_2

continued

Learning Activity 4.4 (continued)

2. In each of the following questions, identify the *limiting factor* and the *excess reactant*, and then calculate the *amount of water formed* from the amounts given. Assume STP conditions. Here is the balanced chemical equation:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

   a) 0.500 moles of H_2 and 0.750 moles of O_2
   b) 0.800 moles H_2 and 0.750 moles O_2
   c) 5.00 g H_2 and 56.0 g O_2
   d) 2.00 L H_2 gas and 2.00 L O_2 gas at STP
   e) 7.00 L H_2 gas and 3.00 L O_2 gas at STP

3. Given 0.161 g of hydrogen gas and 5.62 g of nitrogen gas, calculate the mass of HN_3 produced. Calculate the mass of excess reactant that remains. Use the following balanced equation:

$$H_2 + 3N_2 \rightarrow 2HN_3$$

4. According to the unbalanced reaction below,

$$AlBr_3 + Cl_2 \rightarrow Br_2 + AlCl_3$$

   how many grams of aluminum chloride are produced from 82.5 g of chlorine and 175 g of aluminum bromide? How many grams of the excess reactant remain?

5. For the unbalanced reaction,

$$Al_{(s)} + Br_{2(g)} \rightarrow AlBr_{3(s)}$$

   what mass of aluminum bromide can be made from 50.0 g of aluminum and 50.0 L of bromine at STP?

6. For the unbalanced reaction,

$$Fe_2O_{3(s)} + CO_{(g)} \rightarrow Fe_3O_{4(s)} + CO_{2(g)}$$

   what volume of carbon dioxide is formed from 102 g of iron (III) oxide and 6.50 L of carbon monoxide at STP?

Lesson Summary

In this lesson, you learned that the limiting reactant is the reactant completely consumed in a reaction, while the excess reactant is the reactant remaining after the limiting reactant is used up. Knowing that an insufficient quantity of any of the reactants will limit the amount of product formed in a reaction, you practiced the steps involved in solving limiting reactant problems. In the next lesson, you will complete a hands-on investigation to identify a limiting reactant and calculate the mole ratio.

NOTES

Assignment 4.4

Solving Limiting Reactant Problems *(24 marks)*

In each of the following questions, identify the *limiting factor* and the *excess reactant*, and then calculate the *amount of the indicated product* formed from the amounts given.

1. Given 3.0 moles of methane and 4.0 moles of oxygen gas, calculate the moles of carbon dioxide gas produced. *(5 marks)*

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

2. Given 5.0 mol of acetylene and 11.0 mol of oxygen gas, calculate the volume of CO_2 gas produced. *(6 marks)*

   $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$

continued

Assignment 4.4: Solving Limiting Reactant Problems (continued)

3. Given 160.5 g of sulfur and 268.8 g of oxygen gas, calculate the mass of SO_3 gas produced. How many grams of the excess reactant remain? *(13 marks)*

   $2S + 3O_2 \rightarrow 2SO_3$

Lesson 5: Laboratory Experiment—Limiting Reactant Investigation (2 hours)

Lesson Focus

SLO C11-3-15: Perform a lab involving mass-mass or mass-volume relations, identifying the limiting reactant and calculating the mole ratio.
Include: theoretical yield, experimental yield

Percent Yield

Imagine that on your last Chemistry test you scored an 81/100, or 81%. What does this really mean? Your grade is really a ratio of two numbers, one which represents the number of questions you could have answered correctly (100 possible) and the ones you actually did answer correctly (81). In Chemistry, you can also use this type of ratio to determine how much product is actually formed, compared to what you expected to form, based on the balanced chemical equation.

Before you begin any lab experiment, you can use the balanced chemical equation to calculate the amount of product that should form during the reaction. This is called the **theoretical yield**. This is the maximum amount of product you could possibly form, but often due to small errors in measurement, impure reactants, and other factors, this yield is rarely achieved. Thus, chemists perform another calculation to determine the amount of product that is actually formed during a chemical reaction, called the **actual yield**. The actual yield is almost always less than the theoretical yield. This is an experimental value, meaning you can only obtain this information after completing the reaction.

Using these two calculated values, chemists determine the **percent yield**. The percent yield is the ratio of the actual yield to the theoretical yield and is expressed as a percentage, much like the mark on your recent chemistry test.

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Since there are almost always sources of error in every experiment, the actual yield of a product is usually less than its theoretical yield. As such, the percent yield is rarely 100% and should not normally be larger than 100%.

Materials

In this investigation, you will illustrate the concept of a limiting reactant, using common everyday items. To complete this investigation, you will need the following materials:

- Measuring cup
- Empty 35 mm film canister*
- Antacid tablets (such as "original" Alka-Seltzer®)
- Vinegar
- Balloons
- Ruler

Note

* You may substitute a test tube for the film canister if you have access to one. An empty prescription pill bottle will also work well.

Learning Activity 4.5

Pre-Lab Inquiry

How do you think the volume of gas will change as you increase the number of Alka-Seltzer® tablets you use with each trial of your investigation?

Procedure

A special type of double replacement reaction called neutralization occurs when the sodium bicarbonate ($NaHCO_3$) from an Alka-Seltzer® tablet is added to acetic acid (CH_3COOH). The products of this reaction are sodium acetate, $NaCH_3COO$, and carbonic acid. Since carbonic acid (H_2CO_3) is very unstable, it will decompose to form water and carbon dioxide gas. In this investigation, you will attempt to collect the carbon dioxide gas as it forms, and estimate the volume of gas produced from each trial. The balanced chemical equation is as follows:

$$NaHCO_3 + CH_3COOH \rightarrow NaCH_3COO + H_2O + CO_2$$

continued

Learning Activity 4.5 (continued)

Here are the steps to follow and complete for this investigation. You will complete three trials using the same series of steps each time. What will change with each trial is the amount of Alka-Seltzer® you add to the same volume of vinegar (acetic acid). It is advisable to read these steps over before you begin. If you need help getting started, contact your tutor/marker with your questions.

1. Add enough vinegar to fill the film canister about half full.
2. Break one Alka-Seltzer® tablet in half and then in half again. Add about one-quarter of an Alka-Seltzer® tablet to the vinegar.
3. Quickly seal the balloon over the top of the film canister.
4. As the vinegar reacts with the Alka-Seltzer®, a certain amount of gas will form and fill the film canister. To help the process along, pick up the canister and gently swirl the contents. Don't let the balloon separate from the canister!
5. When it appears that no more gas is produced, tie off the balloon, and then use your ruler to estimate the volume of gas in the balloon. First, find the widest point on the balloon and estimate its diameter. Divide the value by two to obtain the radius. Remember that to calculate the volume of a sphere, you must use the following formula:

$$\text{volume} = \frac{4}{3}\pi r^3$$

6. Rinse the canister and try to remove as much water as possible.
7. Repeat the above steps for trial 2, this time using the same amount of vinegar and one-half of an Alka-Seltzer® tablet.
8. Repeat the above steps for trial 3, this time using the same amount of vinegar and one full Alka-Seltzer® tablet.
9. Complete the analysis and conclusion questions.

Lesson Summary

In this lesson, you applied your knowledge about limiting reactants by completing a simple at-home experiment. In the next lesson, you will continue learning about applications of stoichiometry, with a focus on industry.

Assignment 4.5

Limiting Reactant Investigation *(22 marks)*

1. Complete the following results table. Show your volume and mole calculations in the space provided. Because you are working at room temperature, assume conditions of SATP, where molar volume is 24.4 L/mol. See Lesson 7 from Module 3 if you need to review calculations involving molar volume. Note that 1 cm^3 = 1 mL. Each calculation is worth one mark each.

   Calculations: (9 marks)

Trial	Alka-Seltzer® (portion of tablet)	Volume of gas produced (cm^3)	Volume of gas produced (L)	Moles of gas produced (mol)
1	0.25			
2	0.50			
3	1.00			

2. How did the volumes of gas produced, as measured by the size of the balloon, compare? *(1 mark)*

continued

3. Show by calculation why the balloon in Trial 2 inflated to about twice the size of the balloon in Trial 1. *(2 marks)*

4. What was the limiting reactant in each trial of this experiment? Explain your answer. *(2 marks)*

5. According to the packaging, each complete Alka-Seltzer® tablet contains 1.916 grams of sodium bicarbonate. Calculate:
   a) The moles of sodium bicarbonate. *(2 marks)*

   b) The moles of carbon dioxide that could theoretically be produced by the tablet. *(2 marks)*

continued

Assignment 4.5: Limiting Reactant Investigation (continued)

c) The volume of carbon dioxide that could theoretically be produced at SATP. *(2 marks)*

6. Calculate the percent yield in your investigation by comparing your theoretical volume from 5(c) to your actual volume from Trial 3 in your data table. *(2 marks)*

NOTES

Lesson 6: Stoichiometry Applications
(2 hours)

Lesson Focus

SLO C11-3-16: Discuss the importance of stoichiometry in industry and describe specific applications.
Examples: analytical chemistry, chemical engineering, industrial chemistry...

Stoichiometry and Chemical Engineering

Chemically engineered commodities such as soap, tires, aspirin, deodorant, and chocolate bars all rely on stoichiometry for their production.

Baking

When baking, we use recipes to tell us the exact amount of each ingredient required. If you require more of a particular recipe, you can simply double or triple the amounts listed for each ingredient. The same principle applies to stoichiometric constants. Once your equation is balanced, you can "double the recipe" by multiplying all coefficients by two.

Supplements and Fertilizers

Reactions that occur in an organism (i.e., in your body or in a plant) require specific reactants. For example, magnesium (Mg) and manganese (Mn) are important for DNA synthesis. If an organism is deficient in a particular nutrient, this limiting reactant will slow down and even stop cellular growth. Likewise, plants require certain nutrients to grow. To ensure that plants receive adequate amounts of the essential nutrients (such as nitrogen, potassium, phosphorus, calcium, sulfur, and magnesium) farmers treat crops with fertilizers. Consider this practice similar to humans ingesting vitamin supplements to make sure that they receive the daily recommended amount of nutrients.

Perhaps you have noticed fertilizer being distributed on crops while driving by on a nearby highway. Both wet and dry varieties of fertilizer are sprayed from a truck that carries a reservoir. You may know from a previous science course that plants cannot obtain nitrogen directly from their environment. Instead, they must depend on a process called nitrogen fixation, whereby symbiotic bacteria on the root nodules convert nitrogen in the soil to ammonia (NH_3), the ammonium ion (NH_4^+), or the nitrate ion (NO_3^-). These compounds are important for the growth of plants.

A **fertilizer** refers to any compound that contains one or more chemical elements—organic or inorganic, natural or synthetic. The fertilizer can be placed on or mixed into the soil, or it can be directly applied to plants to achieve normal growth. The main types of plant nutrients include organic manures, plant residues, biological nitrogen fixation, and commercial inorganic fertilizers. Chemical fertilizers refer to commercially manufactured products containing a substantial amount of one or more plant nutrients. One such chemical variety is nitrogen fertilizers, which contain ammonia or ammonium ion salts. Ammonia is prepared by the reaction between nitrogen and hydrogen:

$$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$$

Ammonium sulfate (one example of an ammonium ion salt) is made in two steps:

$$2NH_{3(aq)} + CO_{2(aq)} + H_2O_{(l)} \rightarrow (NH_4)_2CO_{3(aq)}$$

$$(NH_4)_2CO_{3(aq)} + CaSO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)} + CaCO_{3(s)}$$

To increase the yield in the previous two reactions, ammonia is made the limiting reactant in the first step, while ammonium carbonate is the limiting reactant in the second step.

Stoichiometry and Industry

Most industrial processes that involve chemical reactions rely heavily on stoichiometry to help determine how much of a reactant must be purchased. In industry, stoichiometry is used to “scale up” chemical reactions to produce very large quantities of high quality products at a low cost. Considerations in this process include efficient engineering, increased profitability, reduced waste, and heightened safety (i.e., large amounts of heat produced by large-scale chemical reactions must be carefully managed to prevent explosions).

Stoichiometric relations between acids and bases are used to determine unknown concentrations or amounts through a process called titration (a carefully controlled neutralization reaction). For example, an environmental chemist studying a lake in which fish are dying can perform a neutralization reaction to find out how much acid is present in a sample of water from the lake.

Automobiles

Catalytic Converters

When air is drawn into a car’s engine, both nitrogen and oxygen are present. When gasoline is burned in the engine to produce the energy that drives your car, nitrogen and oxygen sometimes combine to form nitrogen monoxide (NO). If the nitrogen monoxide gets into the atmosphere, it gets oxidized to nitrogen dioxide (NO_2). You would recognize nitrogen dioxide as the brownish haze you see over cities that suffer from severe air pollution, like New York and Los Angeles. Nitrogen dioxide can cause additional damage in the atmosphere by reacting with water to produce nitric acid (a component of acid rain) or by decomposing to form an oxygen atom that combines with molecular oxygen to form damaging ozone molecules near the Earth’s surface.

To reduce the amount of air pollution from automobiles, catalysts (e.g., platinum, palladium, or rhodium metals) in your car’s catalytic converter serve as a surface to which nitrogen monoxide molecules stick, and then decompose. Because the nitrogen and oxygen don’t remain on the catalyst’s surface, the catalyst returns to its original state after the reaction and can be reused. The nitrogen and oxygen end up in the carbon dioxide and water exhaust gases.

Unfortunately, catalytic converters aren’t perfect. Pollutants are still emitted when you start your car, because the converters don’t work below 316 °C, and it takes some time to reach that temperature as you drive your car.

Oxygen Sensors

In order to decrease emissions, modern car engines with fuel injection systems control the amount of fuel that is burned with the help of an oxygen sensor. This sensor, placed between the engine and the catalytic converter, tells the engine computer how much oxygen is in the exhaust. The computer can then change the fuel to air ratio to make sure that the engine is running at the proper stoichiometric ratio.

Air Bags

Depending on the make and model year of your vehicle, you may have one of several air bags that contain a small amount of a compound called sodium azide. If your car is involved in a collision, a motion sensor releases a spark that causes an immediate and explosive decomposition reaction:

$$2NaN_{3(s)} \rightarrow 2Na_{(s)} + 3N_{2(g)}$$

The result of the reaction is the production of nitrogen gas, which inflates the bag one-tenth of a second after impact. The inflated airbag(s) protect you and your passenger from entering into a second collision with the steering wheel, windshield, or other component of your car. Here, stoichiometry is used to ensure that air bags do not underinflate or overinflate under varying conditions of temperature. Here are two additional reactions occurring within the air bag:

$$6Na_{(s)} + Fe_2O_{3(s)} \rightarrow 3Na_2O_{(s)} + 2Fe_{(s)} + 418\ kJ$$

and

$$Na_2O_{(s)} + 2CO_{2(g)} + H_2O_{(g)} \rightarrow 2NaHCO_{3(s)}$$

Breath Analysis

The traditional breathalyzer measures the amount of alcohol in the breath using the reaction:

$$3CH_3CH_2OH + 2Cr_2O_7^{2-} + 16H+ \rightarrow 4Cr^{3+} + 3CH_3COOH + 11H_2O$$

In years past, a breath sample suspected of containing ethanol was bubbled through a glass cylinder containing acidified potassium dichromate to determine the percent alcohol in one’s blood. A change from the yellow/orange colour of the dichromate to the green of the chromium (III) ion was directly proportional to the concentration of ethanol (the limiting reactant) in the blood. Thus, the greater the amount of ethanol in the blood, the darker the chromium (III) ion product. Today, lasers are used to determine blood alcohol content more quickly and accurately.

Lesson Summary

In this lesson, you discovered some of the ways stoichiometry is used in industry. You learned that applications of stoichiometry include airbags in cars, fertilizers in agriculture, and breathalyzer tests.

NOTES

Assignment 4.6

Stoichiometry Applications *(23 marks)*

The Stoichiometry of Gasoline

The energy used to propel your vehicle comes from the combustion of gasoline (C_8H_{18}). Here is the balanced equation:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

In this assignment, you will investigate the stoichiometry of gasoline using three different vehicles as examples.

1. Complete the following table. (If you prefer to substitute these three examples with your choice of cars, feel free to indicate this on the table). Use the most current model available, as well as Canadian references to ensure that values are given in the proper units (e.g., Nissan Canada). If you can't find this information, contact your tutor/marker for assistance. *(3 marks)*

Car	Honda Civic (four door sedan)	Lexus SC430	Nissan Pathfinder SE
Size of gas tank (L)			
City fuel economy (L/100 km)			
Highway fuel economy (L/100 km)			

continued

Assignment 4.6: Stoichiometry Applications (continued)

2. What is the cost of:
   a) Regular gasoline? ______________ *(1 mark)*
   b) Premium/supreme gasoline? ______________ *(1 mark)*
   c) Calculate the cost to fill the gas tank of each vehicle. Note that the Lexus requires premium gasoline. Show your calculations. *(3 marks)*

      Honda Civic: ____

      Lexus SC430: ____

      Pathfinder SE: ____

3. If one litre of gasoline contains 6.18 moles of gas, how many moles fit into the tank of each vehicle? Show your calculations. *(3 marks)*

   Honda Civic: ____

   Lexus SC430: ____

   Pathfinder SE: ____

4. How many moles, grams, and litres of carbon dioxide gas would be emitted from each vehicle if you were to use an entire tank of gasoline in each car? Show each calculation that you perform. *(9 marks)*

	Honda Civic (four door sedan)	**Lexus SC430**	**Nissan Pathfinder SE**
Moles of carbon dioxide			
Kilograms of carbon dioxide			
Litres of carbon dioxide			

continued

Assignment 4.6: Stoichiometry Applications (continued)

5. If you consumed one tank of gasoline every two weeks, what would be the volume of carbon dioxide gas emitted by each car in one year? Show your work. *(3 marks)*

   Honda Civic: ______________________________

   Lexus SC430: ______________________________

   Pathfinder SE: ______________________________

NOTES

Module 4 Summary

Congratulations! You have reached the end of Module 4.

Submitting Your Assignments

It is now time for you to submit Assignments 4.1 to 4.6 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 4 assignments and organize your material in the following order:

- Module 4 Cover Sheet (found at the end of the course Introduction)
- Assignment 4.1: Interpreting a Balanced Equation
- Assignment 4.2: Using the Molar Ratio
- Assignment 4.3: Converting between Volume and Mass
- Assignment 4.4: Solving Limiting Reactant Problems
- Assignment 4.5: Limiting Reactant Investigation
- Assignment 4.6: Stoichiometry Applications

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

NOTES

Grade 11 Chemistry (30S)

Module 4: Stoichiometry

Learning Activity Answer Keys

Module 4: Stoichiometry

Learning Activity 4.1: Describing a Balanced Equation

1. How is a balanced equation similar to a recipe?

   Answer:

   Both a balanced equation and a recipe give information regarding the quantity, or amount, of materials needed to make a final product.

2. Chemical reactions can be described in terms of what quantities?

   Answer:

   Chemical reactions can be described in terms of the following quantities: mass, atoms, molecules, formula units, moles, and volumes.

3. Interpret

   $2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$

   in terms of

   a) numbers of moles

      Answer:

      $2 \text{ mol } C_2H_2 + 5 \text{ mol } O_2 \rightarrow 4 \text{ mol } CO_2 + 2 \text{ mol } H_2O$

   b) mass

      Answer:

      (2 moles)(26.0 g/mol) + (5 moles)(32.0 g/mol) → (4 moles)(44.0 g/mol) + (2 moles)(18.0 g/mol)

      $52.0 \text{ g } C_2H_2 + 160.0 \text{ g } O_2 \rightarrow 176.0 \text{ g } CO_2 + 36.0 \text{ g } H_2O$

      212.0 g reactants → 212.0 g products

   c) volumes of gases at STP

      Answer:

      (2 moles)(22.4 L/mol) + (5 moles)(22.4 L/mol) → (4 moles)(22.4 L/mol) + (2 moles)(22.4 L/mol)

      $44.8 \text{ L } C_2H_2 + 112 \text{ L } O_2 \rightarrow 89.6 \text{ L } CO_2 + 44.8 \text{ L } H_2O$

Learning Activity 4.2: Solving Stoichiometric Problems

Show your work in solving these stoichiometric problems, including the ratios you used. Try to determine the appropriate number of significant figures for your final answer. Don't forget to balance the equation first! Remember that molar masses have all been rounded and are NOT included in your significant figure determination.

Use the following unbalanced reaction to answer questions 1 to 3.

$NH_3 + O_2 \rightarrow N_2 + H_2O$

1. a) What are the molar coefficients of the balanced reaction?

Answer:

4, 3, 2, 6

b) How many moles of oxygen gas will react exactly with 1.60 mol of ammonia?

Answer:

$$1.60\ \cancel{\text{mol NH}_3} \times \frac{3\ \text{mol O}_2}{4\ \cancel{\text{mol NH}_3}} = 1.20\ \text{mol O}_2$$

c) How many moles of each product will be generated by 1.60 mol of ammonia?

Answer:

$$1.60\ \cancel{\text{mol NH}_3} \times \frac{2\ \text{mol N}_2}{4\ \cancel{\text{mol NH}_3}} = 0.800\ \text{mol N}_2$$

$$1.60\ \cancel{\text{mol NH}_3} \times \frac{6\ \text{mol H}_2\text{O}}{4\ \cancel{\text{mol NH}_3}} = 2.40\ \text{mol H}_2\text{O}$$

2. a) How many moles of oxygen gas will react with 0.750 mol of ammonia gas?

Answer:

$$0.750\ \cancel{\text{mol NH}_3} \times \frac{3\ \text{mol O}_2}{4\ \cancel{\text{mol NH}_3}} = 0.5625\ \text{mol} \approx 0.562\ \text{mol O}_2$$

b) How many moles of each product will be produced from the 0.750 mol of ammonia?

Answer:

$$0.750\ \cancel{\text{mol NH}_3} \times \frac{2\ \text{mol N}_2}{4\ \cancel{\text{mol NH}_3}} = 0.375\ \text{mol N}_2$$

$$0.750\ \cancel{\text{mol NH}_3} \times \frac{6\text{mol H}_2\text{O}}{4\ \cancel{\text{mol NH}_3}} = 1.125\ \text{mol} \approx 1.12\ \text{mol H}_2\text{O}$$

3. Determine the number of moles of water that would be produced from 2.50 mol of ammonia reacting with an excess of oxygen gas.

Answer:

$$2.50\ \cancel{\text{mol NH}_3} \times \frac{6\ \text{mol H}_2\text{O}}{4\ \cancel{\text{mol NH}_3}} = 3.75\ \text{mol H}_2\text{O}$$

4. How many moles of H_2S can be burned by 0.750 moles of oxygen gas?

$$2H_2S + 3O_2 \rightarrow 2H_2O + 2SO_2$$

Answer:

$$0.750\ \cancel{\text{mol O}_2} \times \frac{2\ \text{mol H}_2\text{S}}{3\ \cancel{\text{mol O}_2}} = 0.500\ \text{mol H}_2\text{S}$$

5. How many moles of oxygen can be produced from 1.80 moles of $KClO_3$?

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

Answer:

$$1.80\ \cancel{\text{mol KClO}_3} \times \frac{3\ \text{mol O}_2}{2\ \cancel{\text{mol KClO}_3}} = 2.70\ \text{mol O}_2$$

6. Calculate the mass of oxygen and hydrogen that will be formed by the decomposition of 4.50 g of water, according to the following reaction.

   $2H_2O \rightarrow 2H_2 + O_2$

   $H_2O = 18.0$ g/mol, $H_2 = 2.0$ g/mol, $O_2 = 32.0$ g/mol

   Answer:

$$4.50\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.0\ \cancel{\text{g H}_2\text{O}}} \times \frac{1\ \cancel{\text{mol O}_2}}{2\ \cancel{\text{mol H}_2\text{O}}} \times \frac{32.0\ \text{g O}_2}{1\ \cancel{\text{mol O}_2}} = 4.00\ \text{g O}_2$$

$$4.50\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.0\ \cancel{\text{g H}_2\text{O}}} \times \frac{2\ \cancel{\text{mol H}_2}}{2\ \cancel{\text{mol H}_2\text{O}}} \times \frac{2.0\ \text{g H}_2}{1\ \cancel{\text{mol H}_2}} = 0.500\ \text{g H}_2$$

7. Calculate the mass of aluminum oxide produced from 8.00 g of oxygen gas reacting with an excess of aluminum metal, according to the following reaction.

   $4Al + 3O_2 \rightarrow 2Al_2O_3$

   $O_2 = 32.0$ g/mol, $Al_2O_3 = 102.0$ g/mol

   Answer:

$$8.00\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.0\ \cancel{\text{g O}_2}} \times \frac{2\ \cancel{\text{mol Al}_2\text{O}_3}}{3\ \cancel{\text{mol O}_2}} \times \frac{102.0\ \text{g Al}_2\text{O}_3}{1\ \cancel{\text{mol Al}_2\text{O}_3}} = 17.0\ \text{g Al}_2\text{O}_3$$

8. What mass of water is needed to react exactly with 2.30 g of NO_2 gas, and what mass of HNO_3 will be formed, according to the following reaction?

   $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

   $NO_2 = 46.0$ g/mol, $H_2O = 18.0$ g/mol, $HNO_3 = 63.0$ g/mol

   Answer:

$$2.30\ \cancel{\text{g NO}_2} \times \frac{1\ \cancel{\text{mol NO}_2}}{46.0\ \cancel{\text{g NO}_2}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{3\ \cancel{\text{mol NO}_2}} \times \frac{18.0\ \text{g H}_2\text{O}}{1\ \cancel{\text{mol H}_2\text{O}}}$$

$$= 0.300\ \text{g H}_2\text{O}$$

$$2.30\ \cancel{\text{g NO}_2} \times \frac{1\ \cancel{\text{mol NO}_2}}{46.0\ \cancel{\text{g NO}_2}} \times \frac{2\ \cancel{\text{mol HNO}_3}}{3\ \cancel{\text{mol NO}_2}} \times \frac{63.0\ \text{g HNO}_3}{1\ \cancel{\text{mol HNO}_3}}$$

$$= 2.10\ \text{g HNO}_3$$

9. What mass of tin (II) nitrate will be formed from the reaction of 25.2 g of nitric acid (HNO_3) and an excess of tin, according to the following reaction?

$4Sn + 10HNO_3 \rightarrow 4Sn(NO_3)_2 + NH_4NO_3 + 3H_2O$

HNO_3 = 63.0 g/mol, $Sn(NO_3)_2$ = 242.7 g/mol

Answer:

$$25.2\ \cancel{\text{g HNO}_3} \times \frac{1\ \cancel{\text{mol HNO}_3}}{63.0\ \cancel{\text{g HNO}_3}} \times \frac{4\ \text{mol}\ \cancel{\text{Sn(NO}_3)_2}}{10\ \cancel{\text{mol HNO}_3}} \times \frac{242.7\ \text{g Sn(NO}_3)_2}{1\ \text{mol}\ \cancel{\text{Sn(NO}_3)_2}}$$

$$= 38.8\ \text{g Sn(NO}_3)_2$$

10. What mass of HCl is required to form 14.2 g of Cl_2?

$4HCl + O_2 \rightarrow 2H_2O + 2Cl_2$

HCl = 36.5 g/mol; Cl_2 = 71.0 g/mol

Answer:

$$14.2\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{71.0\ \cancel{\text{g Cl}_2}} \times \frac{4\ \cancel{\text{mol HCl}}}{2\ \cancel{\text{mol Cl}_2}} \times \frac{36.5\ \text{g HCl}}{1\ \cancel{\text{mol HCl}}} = 14.6\ \text{g HCl}$$

11. What mass of H_3PO_4 will react with 60.0 g of NaOH ?

$H_3PO4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$

NaOH = 40.0 g/mol, H_3PO_4 = 98.0 g/mol

Answer:

$$60.0\ \cancel{\text{g NaOH}} \times \frac{1\ \cancel{\text{mol NaOH}}}{40.0\ \cancel{\text{g NaOH}}} \times \frac{1\ \cancel{\text{mol H}_3\text{PO}_4}}{3\ \cancel{\text{mol NaOH}}} \times \frac{98.0\ \text{g H}_3\text{PO}_4}{1\ \cancel{\text{mol H}_3\text{PO}_4}}$$

$$= 49.0\ \text{g H}_3\text{PO}_4$$

12. How many moles of hydrogen gas are formed from 18.3 g of HCl ?

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

HCl = 36.5 g/mol, H_2 = 2.0 g/mol

Answer:

$$18.3\ \cancel{\text{g HCl}} \times \frac{1\ \cancel{\text{mol HCl}}}{36.5\ \cancel{\text{g HCl}}} \times \frac{1\ \cancel{\text{mol H}_2}}{2\ \cancel{\text{mol HCl}}} \times \frac{2.0\ \text{g H}_2}{1\ \cancel{\text{mol H}_2}} = 0.501\ \text{g H}_2$$

Learning Activity 4.3: Using the Molar Ratio to Calculate Volume

Show your work in solving these stoichiometric problems, including the ratios you used in your calculation. Balance the chemical equation when necessary and practice using the appropriate number of significant figures in your final answer.

1. What volume of each product could be made from 8.00 g of methane gas (CH_4) at STP, according to the following reaction?

   $$CH_4 \rightarrow C_2H_2 + H_2$$

   Answer:

   $$2CH_4 \rightarrow C_2H_2 + 3H_2$$

   $CH_4 = 16.0$ g/mol

   $$V_{C_2H_2} = 8.00\ \cancel{g}\ \cancel{CH_4} \times \frac{1\ \cancel{mol}\ \cancel{CH_4}}{16.0\ \cancel{g}\ \cancel{CH_4}} \times \frac{1\ \cancel{mol}\ \cancel{C_2H_2}}{2\ \cancel{mol}\ \cancel{CH_4}} \times \frac{22.4\ L\ C_2H_2}{1\ \cancel{mol}\ \cancel{C_2H_2}}$$

   $$= 5.60\ L\ C_2H_2$$

   or

   $$V_{C_2H_2} = \frac{8.00\ \cancel{g}\ \cancel{CH_4}}{16.0\ \cancel{g}/\cancel{mol}} \times \frac{1\ \cancel{mol}\ C_2H_2}{1\ \cancel{mol}\ \cancel{CH_4}} \times \frac{22.4\ L}{1\ \cancel{mol}} = 5.60\ L\ C_2H_2$$

   $$V_{H_2} = 8.00\ g\ \cancel{CH_4} \times \frac{1\ \cancel{mol}\ \cancel{CH_4}}{16.0\ g\ \cancel{CH_4}} \times \frac{3\ \cancel{mol\ H_2}}{2\ \cancel{mol}\ \cancel{CH_4}} \times \frac{22.4\ L\ H_2}{1\ \cancel{mol\ H_2}}$$

   $$= 16.8\ L\ H_2$$

   or

   $$V_{H_2} = \frac{8.00\ \cancel{g}\ \cancel{CH_4}}{16.0\ \cancel{g}\ /\ \cancel{mol}} \times \frac{3\ \cancel{mol}\ H_2}{2\ \cancel{mol}\ \cancel{CH_4}} \times \frac{22.4\ L}{1\ \cancel{mol}} = 16.8\ L\ H_2$$

2. What volume of CO_2 gas can be made from 11.2 L of CO gas and an excess of iron (III) oxide? Assume that temperature and pressure are kept constant at STP.

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

Answer:

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

$$V_{CO_2} = 11.2\ \cancel{L\ CO} \times \frac{1\ \cancel{mol\ CO}}{22.4\ \cancel{L\ CO}} \times \frac{3\ \cancel{mol\ CO_2}}{3\ \cancel{mol\ CO}} \times \frac{22.4\ L\ CO_2}{1\ \cancel{mol\ CO_2}}$$

$$= 11.2\ L\ CO_2$$

or

$$V_{CO_2} = \frac{11.2\ \cancel{L}\ \cancel{CO}}{22.4\ \cancel{L}/\cancel{mol}} \times \frac{3\ \cancel{mol}\ CO_2}{3\ \cancel{mol}\ \cancel{CO}} \times \frac{22.4\ L}{1\ \cancel{mol}} = 11.2\ L\ CO_2$$

3. Given:

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

How many litres of nitrogen dioxide are produced when 34.0 L of oxygen gas reacts with an excess of nitrogen monoxide at STP?

Answer:

$$V_{NO_2} = 34.0\ \cancel{L\ O_2} \times \frac{1\ \cancel{mol\ O_2}}{22.4\ \cancel{L\ O_2}} \times \frac{2\ \cancel{mol\ NO_2}}{1\ \cancel{mol\ O_2}} \times \frac{22.4\ L\ NO_2}{1\ \cancel{mol\ NO_2}}$$

$$= 68.0\ L\ NO_2$$

or

$$V_{NO_2} = \frac{34.0\ \cancel{L}\ \cancel{O_2}}{22.4\ \cancel{L}/\cancel{mol}\ \cancel{NO_2}} \times \frac{2\ \cancel{mol}\ \cancel{NO_2}}{1\ \cancel{mol}\ \cancel{O_2}} \times \frac{22.4\ L}{1\ \cancel{mol}} = 68.0\ L\ NO_2$$

4. $C_3H_8O_2 + 4O_2 \rightarrow 3CO_2 + 4H_2O$

In the reaction above, the heat of combustion is 420.0 kJ/mole of $C_3H_8O_2$. When 125 L of $C_3H_8O_2$ react (at STP), how much energy would be released?

Answer:

$$125\ \cancel{L} \times \frac{1\ \text{mol}}{22.4\ \cancel{L}} = 5.58\ \text{mol}$$

$$5.58\ \cancel{\text{mol}} \times \frac{420.0\ \text{kJ}}{1\ \cancel{\text{mol}}} = 2343.75\ \text{kJ} = 2340\ \text{kJ} = 2.34 \times 10^3\ \text{kJ}$$

or

$$125\ \cancel{L} \times \left(\frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{L}}\right) \times \left(\frac{420.0\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) = 2343.75\ \text{kJ} = 2340\ \text{kJ} = 2.34 \times 10^3\ \text{kJ}$$

Learning Activity 4.4: Limiting Factor and Excess Reactant

1. In each of the following questions, identify the *limiting factor* and the *excess reactant*. Use the following balanced equation:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

a) 2 molecules of H_2 and 2 molecules of O_2

Answer:

2 molecules of H_2 react with 1 molecule of O_2. 4 molecules of H_2 are needed to react with 2 molecules of O_2. Since there are only 2 molecules of H_2, the H_2 is the limiting factor and O_2 is the excess reactant.

b) 10 molecules of H_2 and 4 molecules of O_2

Answer:

$$10\ \cancel{\text{molecules } H_2} \times \frac{1\ \text{molecules } O_2}{2\ \cancel{\text{molecules } H_2}} = 5\ \text{molecules } O_2$$

Only 4 molecules of O_2 are given. Therefore, O_2 is the limiting factor and H_2 is the excess reactant.

c) 50 molecules of H_2 and 20 molecules of O_2

Answer:

$$50\ \cancel{\text{molecules } H_2} \times \frac{1\ \text{molecules } O_2}{2\ \cancel{\text{molecules } H_2}} = 25\ \text{molecules } O_2$$

O_2 is the limiting factor and H_2 is the excess reactant.

d) 100 molecules of H_2 and 75 molecules of O_2

Answer:

$$100\ \cancel{\text{molecules H}_2} \times \frac{1\ \text{molecules O}_2}{2\ \cancel{\text{molecules H}_2}} = 50\ \text{molecules O}_2$$

More than 50 molecules of O_2 are given. There is not enough H_2 to use up all of the O_2. Therefore, H_2 is the limiting factor and O_2 is the excess reactant.

e) 100 molecules of H_2 and 25 molecules of O_2

Answer:

$$100\ \cancel{\text{molecules H}_2} \times \frac{1\ \text{molecules O}_2}{2\ \cancel{\text{molecules H}_2}} = 50\ \text{molecules O}_2$$

Only 25 molecules of O_2 are given. Therefore, O_2 is the limiting factor and H_2 is the excess reactant.

2. In each of the following questions, identify the limiting factor and the excess reactant, and then calculate the amount of water formed from the amounts given. Assume STP conditions. Here is the balanced chemical equation:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

a) 0.500 moles of H_2 and 0.750 moles of O_2

Answer:

$$0.500\ \cancel{\text{mol H}_2} \times \frac{1\ \text{mol O}_2}{2\ \cancel{\text{mol H}_2}} = 0.250\ \text{mol O}_2$$

H_2 is the limiting factor and O_2 is the excess reactant.

$$\text{Mass of water} = 0.500\ \cancel{\text{mol H}_2} \times \frac{2\ \cancel{\text{mol H}_2\text{O}}}{2\ \cancel{\text{mol H}_2}} \times \frac{18.0\ \text{g H}_2\text{O}}{\cancel{\text{mol H}_2\text{O}}}$$

$$= 9.00\ \text{g H}_2\text{O}$$

b) 0.800 moles H_2 and 0.750 moles O_2

Answer:

$$0.800\ \cancel{\text{mol } H_2} \times \frac{1\ \text{mol } O_2}{2\ \cancel{\text{mol } H_2}} = 0.400\ \text{mol } O_2$$

H_2 is the limiting factor and O_2 is the excess reactant.

$$\text{Mass of water} = 0.800\ \cancel{\text{mol } H_2} \times \frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}} \times \frac{18.0\ \text{g}}{\cancel{\text{mol } H_2O}}$$

$$= 14.4\ \text{g } H_2O$$

c) 5.00 g H_2 and 56.0 g O_2

Answer:

$$\text{moles } H_2 = \frac{5.00\ \cancel{\text{g}}}{2.0\ \cancel{\text{g}}/\text{mol}} = 2.50\ \text{mol } H_2$$

$$\text{moles } O_2 = \frac{56.0\ \cancel{\text{g}}}{32.0\ \cancel{\text{g}}/\text{mol}} = 1.75\ \text{mol } O_2$$

$$2.5\ \cancel{\text{mol } H_2} \times \frac{1\ \text{mol } O_2}{2\ \cancel{\text{mol } H_2}} = 1.2\ \text{mol } O_2$$

H_2 is the limiting factor and O_2 is the excess reactant.

$$\text{Mass of water} = 2.50\ \cancel{\text{mol } H_2} \times \frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}} \times \frac{18.0\ \text{g}}{\cancel{\text{mol } H_2O}}$$

$$= 45.0\ \text{g } H_2O$$

d) 2.00 L H_2 gas and 2.00 L O_2 gas at STP

Answer:

$$\text{moles } H_2 = \frac{2.00\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 0.0893\ \text{mol } H_2$$

$$\text{moles } O_2 = \frac{2.00\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 0.0893\ \text{mol } O_2$$

$$\text{moles } O_2 = 0.0893\ \cancel{\text{mol } H_2} \times \frac{1\ \text{mol } O_2}{2\ \cancel{\text{mol } H_2}} = 0.0446\ \text{mol } O_2$$

H_2 is the limiting factor and O_2 is the excess reactant.

$$\text{Mass of water} = 0.0893\ \cancel{\text{mol } H_2} \times \frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}} \times \frac{18.0\ \text{g}}{\cancel{\text{mol } H_2O}}$$

$$= 1.61\ \text{g } H_2O$$

e) 7.00 L H_2 gas and 3.00 L O_2 gas at STP

Answer:

$$\text{moles } H_2 = \frac{7.00\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 0.312\ \text{mol } H_2$$

$$\text{moles } O_2 = \frac{3.00\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 0.134\ \text{mol } O_2$$

$$0.312\ \cancel{\text{mol } H_2} \times \frac{1\ \text{mol } O_2}{2\ \cancel{\text{mol } H_2}} = 0.156\ \text{mol } O_2$$

O_2 is the limiting factor and H_2 is the excess reactant.

$$\text{Mass of water} = 0.134\ \cancel{\text{mol } O_2} \times \frac{2\ \cancel{\text{mol } H_2O}}{1\ \cancel{\text{mol } O_2}} \times \frac{18.0\ \text{g}}{\cancel{\text{mol } H_2O}}$$

$$= 4.82\ \text{g } H_2O$$

3. Given 0.161 g of hydrogen gas and 5.62 g of nitrogen gas, calculate the mass of HN_3 produced. Calculate the mass of excess reactant that remains. Use the following balanced equation:

$$H_2 + 3N_2 \rightarrow 2HN_3$$

Answer:

$$\text{moles } H_2 = \frac{0.161\ \cancel{g}}{2.0\ \cancel{g}/\text{mol}} = 0.0805 \text{ mol } H_2$$

$$\text{moles } N_2 = \frac{5.62\ \cancel{g}}{28.0\ \cancel{g}/\text{mol}} = 0.201 \text{ mol } N_2$$

$$\text{moles } N_2 = 0.0805\ \cancel{\text{mol } H_2} \times \frac{3 \text{ mol } N_2}{1\ \cancel{\text{mol } H_2}} = 0.242 \text{ mol } N_2$$

0.242 moles of N_2 are needed to react with all of the H_2. Since you are only given 0.201 moles of N_2, *N_2 is the limiting factor and H_2 is the excess reactant.*

HN_3 = 43.0 g/mol

$$\text{mass } HN_3 = 0.201\ \cancel{\text{mol } N_2} \times \frac{2\ \cancel{\text{mol}}\ HN_3}{3\ \cancel{\text{mol } N_2}} \times \frac{43.0 \text{ g}}{1\ \cancel{\text{mol}}} = 5.76 \text{ g } HN_3$$

$$\text{moles } H_2 = 0.201\ \cancel{\text{mol } N_2} \times \frac{1 \text{ mol } H_2}{3\ \cancel{\text{mol } N_2}} = 0.0670 \text{ mol } H_2$$

moles H_2 remaining = initial moles – reacted moles
= 0.0805 mol – 0.0670 mol
= 0.0135 mol H_2 excess

$$\text{mass } H_2 \text{ remaining} = 0.0135\ \cancel{\text{mol}} \times \frac{2.0 \text{ g}}{\cancel{\text{mol}}} = 0.0270 \text{ g } H_2 \text{ excess}$$

4. According to the unbalanced reaction below,

$$AlBr_3 + Cl_2 \rightarrow Br_2 + AlCl_3$$

how many grams of aluminum chloride are produced from 82.5 g of chlorine and 175 g of aluminum bromide? How many grams of the excess reactant remain?

Answer:

$2AlBr_3 + 3Cl_2 \rightarrow 3Br_2 + 2AlCl_3$

Cl_2 = 71.0 g/mol $AlBr_3$ = 266.7 g/mol

$$\text{moles } Cl_2 = \frac{82.5\ \cancel{g}}{71.0\ \cancel{g}/\text{mol}} = 1.16 \text{ mol } Cl_2$$

$$\text{moles } AlBr_3 = \frac{175\ \cancel{g}}{266.7\ \cancel{g}/\text{mol}} = 0.656 \text{ mol } AlBr_3$$

$$\text{moles } Cl_2 = 0.656\ \cancel{\text{mol } AlBr_3} \times \frac{3 \text{ mol } Cl_2}{2\ \cancel{\text{mol } AlBr_3}} = 0.984 \text{ mol } Cl_2$$

0.984 mol of Cl_2 is needed to react with all of the $AlBr_3$. You are given 1.16 mol, so the $AlBr_3$ is used up first. *$AlBr_3$ is the limiting factor and Cl_2 is the excess reactant.*

$AlCl_3$ = 133.5 g/mol

$$\text{mass } AlCl_3 = 0.656\ \cancel{\text{mol } AlBr_3} \times \frac{2\ \cancel{\text{mol}}\ AlCl_3}{2\ \cancel{\text{mol } AlBr_3}} \times \frac{133.5 \text{ g}}{1\ \cancel{\text{mol}}} = 87.6 \text{ g } AlCl_3$$

$$\text{moles } Cl_2 = 0.656\ \cancel{\text{mol } AlBr_3} \times \frac{3 \text{ mol } Cl_2}{2\ \cancel{\text{mol } AlBr_3}} = 0.984 \text{ mol } Cl_2$$

moles Cl_2 remaining = initial moles – reacted moles
= 1.16 moles – 0.984 moles
= 0.18 moles C_{l2} excess

$$\text{moles } Cl_2 \text{ remaining} = 0.18\ \cancel{\text{mol}} \times \frac{71.0 \text{ g}}{\cancel{\text{mol}}} = 13 \text{ g } Cl_2 \text{ excess}$$

5. For the unbalanced reaction,

$$Al_{(s)} + Br_{2(g)} \rightarrow AlBr_{3(s)}$$

what mass of aluminum bromide can be made from 50.0 g of aluminum and 50.0 L of bromine at STP?

Answer:

$2Al_{(s)} + 3Br_{2(g)} \rightarrow 2AlBr_{3(s)}$

$$\text{moles } Br_2 = \frac{50.0\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 2.23 \text{ mol } Br_2$$

Al = 27.0 g/mol

$$\text{moles Al} = \frac{50.0\ \cancel{g}}{27.0\ \cancel{g}/\text{mol}} = 1.85 \text{ mol Al}$$

$$\text{moles } Br_2 = 1.85\ \cancel{\text{mol Al}} \times \frac{3 \text{ mol } Br_2}{2\ \cancel{\text{mol Al}}} = 2.78 \text{ mol } Br_2$$

2.78 mol of Br_2 is needed to react with all of the Al. You only have 2.23 mol of Br_2, so *Br_2 is the limiting factor and Al is the excess reactant.*

$AlBr_3$ = 266.7 g/mol

$$\text{mass } AlBr_3 = 2.23\ \cancel{\text{mol } Br_2} \times \frac{2\ \cancel{\text{mol}}\ AlBr_3}{3\ \cancel{\text{mol } Br_2}} \times \frac{266.7 \text{ g}}{1\ \cancel{\text{mol}}} = 396 \text{ g } AlBr_3$$

6. For the unbalanced reaction,

$$Fe_2O_{3(s)} + CO_{(g)} \rightarrow Fe_3O_{4(s)} + CO_{2(g)}$$

what volume of carbon dioxide is formed from 102 g of iron (III) oxide and 6.50 L of carbon monoxide at STP?

Answer:

$$3Fe_2O_{3(s)} + CO_{(g)} \rightarrow 2Fe_3O_{4(s)} + CO_{2(g)}$$

$$\text{moles CO} = \frac{6.50\ \cancel{L}}{22.4\ \cancel{L}/\text{mol}} = 0.290\ \text{mol CO}$$

$3Fe_2O_3 = 159.6\ \text{g/mol}$

$$\text{moles } Fe_2O_3 = \frac{102\ \cancel{g}}{159.6\ \cancel{g}/\text{mol}} = 0.639\ \text{mol } Fe_2O_3$$

$$\text{moles CO} = 0.639\ \cancel{\text{mol } Fe_2O_3} \times \frac{1\ \text{mol CO}}{3\ \cancel{\text{mol } Fe_2O_3}} = 0.213\ \text{mol CO}$$

0.213 moles of CO are needed to react with all of the Fe_2O_3. You are given 0.290 moles of CO, so *Fe_2O_3 is the limiting factor and CO is the excess reactant.*

$$\text{volume } CO_2 = 0.639\ \cancel{\text{mol } Fe_2O_3} \times \frac{1\ \cancel{\text{mol}}\ CO_2}{3\ \cancel{\text{mol } Fe_2O_3}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 4.77\ \text{L } CO_2$$

Notes

Videos that may be used in the next unit:

https://youtu.be/GnZmbmvcY7E

GRADE 11 CHEMISTRY (30S)

Module 5: Solutions

Module 5: Solutions

Introduction

Is hair gel a solution? What about whipping cream? For now, you will have to wait to find out the answer to these questions, as you learn about mixtures, colloids, and the nine types of solutions in this module. Once you have been introduced to the different types of solutions, you will be introduced to the solution process. You may be surprised to know that not every substance dissolves in the same way! You will also discover that solubility is directly affected by temperature.

Next, you will learn about the concentration of a solution and the different ways that concentration can be expressed. You probably already know that drink crystals mixed with very little water form a concentrated solution, while adding lots of water, makes the drink taste diluted.

Toward the end of the module you will have the opportunity to explore the properties of some solutions that you use at home, work, and school.

Assignments in Module 5

When you have completed the assignments for Module 5, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 5.1	Properties of Solutions
3	Assignment 5.2	The Solution Process
4	Assignment 5.3	Solubility of Polar and Non-Polar Substances
5	Assignment 5.4	Interpreting a Solubility Curve
6	Assignment 5.5	Calculating Solubility
8	Assignment 5.6	Lab Activity: The Effect of Salt on the Melting of Ice
9	Assignment 5.7	Calculations Involving Concentration
10	Assignment 5.8	Determining the Concentration of a Solution
11	Assignment 5.9	Solving Dilution Problems

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

LESSON 1: TYPES OF SOLUTIONS (1 HOUR)

Lesson Focus

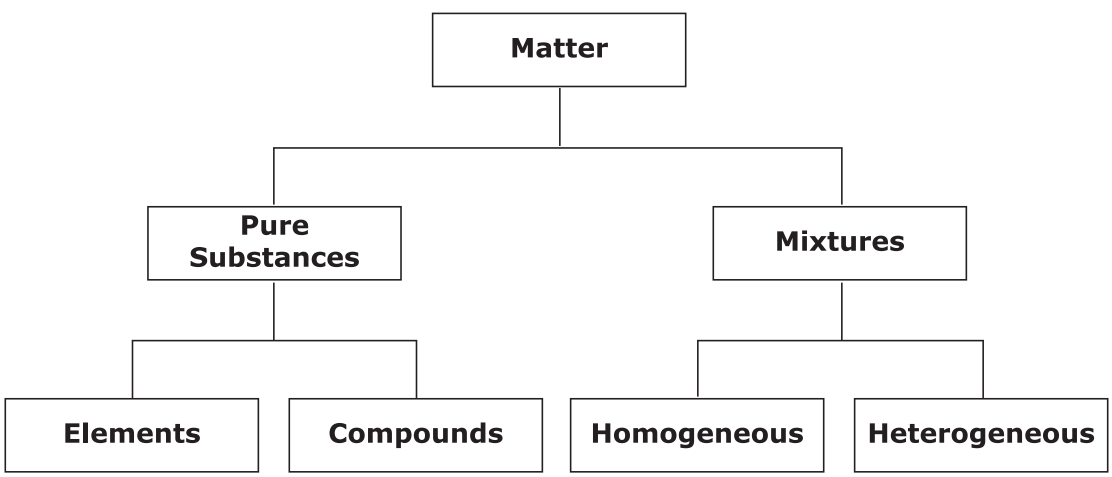
SLO C11-4-01: Describe and give examples of various types of solutions.
Include: all nine possible types

Lesson Introduction

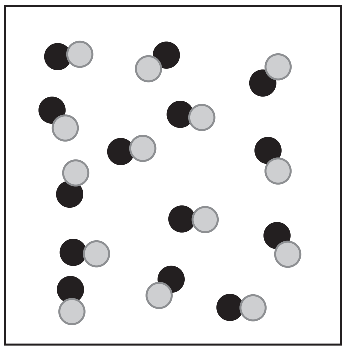
What do Kool-Aid and the air you breathe have in common? Both are examples of solutions that are made up of a dissolving medium (the solvent) and dissolved particles (the solute). In this lesson, you will be introduced to some terminology relating to solutions and learn about the nine different types of solutions.

What is a Mixture?

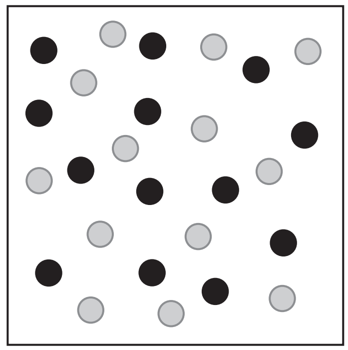
Matter is defined as anything that has mass and volume. You may have already seen a flowchart like this one that illustrated the different kinds of matter:



So far in this course, you have studied the properties of pure substances—substances in which all particles have the same properties. For example, the pure substance shown below could be carbon monoxide, where all the particles are CO (the light circles would represent the C atoms and the dark circles would represent the O atoms), while the mixture could be a mixture of copper particles (represented by the dark circles) and gold particles (represented by the light circles).



Pure Substance



Mixture

Mixtures are different from **pure substances** in that a mixture contains more than one kind of particle. In mixtures, the individual components of the mixture retain their properties. There are two types of mixtures—heterogeneous and homogeneous. Since **solutions** are generally homogeneous mixtures, the lesson will focus mostly on this type of mixture.

Heterogeneous Mixtures

Heterogeneous mixtures will often settle out if left to stand. A common example of a heterogeneous mixture is sand in water. The sand mixes with the water as you swirl the mixture, but eventually settles out on the bottom of the container so that the individual grains of sand are easily distinguished from the water. There are two main types of heterogeneous mixtures—suspensions and colloids.

Suspensions

A **suspension** is different from a solution in that the particles are much larger and will settle out upon standing. You can use a filter to separate the components of a suspension, whereas you cannot accomplish this with a solution.

Colloids

A **colloid** is a mixture of particles that are smaller than those in suspensions, but larger than the particles in solutions. In addition, colloids also have a characteristic cloudy or milky appearance, meaning you cannot see through them. The particles in colloids can be evenly distributed throughout a dispersion medium (either solid, liquid, or gas). Some examples of colloids include glues, whipped cream, paint, gels, emulsions, foams, and aerosols.

Emulsions are another example of colloids, or heterogeneous mixtures, which occur between two or more liquids. An **emulsifying agent** is added to these mixtures to keep the suspension of liquids from separating from each other upon standing. Mayonnaise is a mixture of several liquids, like oil, lemon juice, and eggs. It is a fragile mixture that has the oil suspended as tiny droplets within the lemon juice. The egg yolk acts as the emulsifying agent to keep all of the liquids evenly mixed together. Peanut butter is another example of an emulsion, where the emulsifier lecithin is added to prevent the oil from separating from the rest of the peanut butter mixture. Other examples of emulsions include shampoos, liquid soaps and detergents, and margarine.

Homogeneous Mixtures—Solutions

Homogeneous means to be uniform throughout a sample of material, such that every part of the material is exactly the same as any other part in terms of appearance and phase. Most homogeneous mixtures are known as **solutions** because they are composed of two or more substances that are evenly distributed throughout a single phase. An **aqueous solution** is water that contains dissolved substances. Since water dissolves many of the substances with which it comes in contact, it is difficult to find chemically pure water in nature.

In a solution, the dissolving medium is called the **solvent**, and the dissolved particles are called the **solute**. Usually, a solvent is the component of a solution that is present in a larger amount, while the solute is the component of a solution that is present in a lesser amount. Salt water is an example of a solution where water is the solvent and salt is the solute.

If a substance is able to dissolve in a solvent, the substance is said to be **soluble**. For example, sugar is soluble in water because it dissolves in water. If a substance does not dissolve in a solvent, that substance is said to be **insoluble**. For example, oil is insoluble in water.

Often, insoluble substances are actually soluble in a solvent in very small amounts. However, since the amounts able to dissolve are not significant, these solutions are generally considered to be insoluble. It should be noted, then, that the terms soluble, slightly soluble, and insoluble are relative. This means that even though small amounts of oil can dissolve in water, compared to how much sugar can dissolve in an equal amount of water, a very insignificant amount of oil dissolves.

When two liquids dissolve in each other in any proportion, like alcohol and water, the liquids are said to be **miscible**. When they do not dissolve in any proportion, like oil and water, they are said to be **immiscible**.

Characteristics of Solutions

- Solutions are generally stable mixtures, meaning that the solute will generally not settle out upon standing.
- Solutions are usually clear or transparent, but can still have colour, like cranberry or apple juice. However, solutions do not show what is known as the Tyndall Effect, because if light is shone into a solution from the side, none of the light is reflected to your eye. The solution therefore appears clear, not cloudy. You have likely observed the Tyndall Effect when a ray of sunlight shines through a dusty atmosphere.
- The particles of solute are not distinguishable from the particles of solvent in a solution.
- True solutions cannot be filtered. If we could easily filter saltwater to remove the salt, there would be no shortage of drinking water in the world!

These characteristics differentiate solutions from other mixtures, like colloids, which do not settle out.

Types of Solutions

Most chemical reactions occur in an aqueous medium, and not in the solid, liquid, or gaseous phase. There are nine different types of solutions outlined in the table below. Each type of solution is described in terms of the state (or phase) of the solute and solvent. Remember that the smaller amount in a solution is usually the *solute* while the larger amount is the *solvent*. Can you think of any examples of solutions other than the ones provided?

Solute	Solvent	Examples
Solid	Liquid	Sugar in water Salt in water (saltwater) Iodine in alcohol (tincture)
Liquid	Liquid	Methanol in water (gas line antifreeze) Ethylene glycol in water (engine antifreeze)
Gas	Liquid	Carbon dioxide in water (carbonated beverages) Dissolved oxygen in water (supporting aquatic life)
Solid	Solid	Zinc in copper (brass) 14 karat gold (14 parts gold and 10 parts other metal) Steel Copper in silver (sterling silver)
Liquid	Solid	Mercury in tin or silver (dental amalgams)
Gas	Solid	Hydrogen in palladium Pumice stone
Solid	Gas	Mothballs in air Any substance that is normally solid at room temperature and sublimes Dust in air
Liquid	Gas	Gasoline in air Water vapour in air
Gas	Gas	Air (oxygen, carbon dioxide, etc., in nitrogen)

The solutions using liquids as the solvent are quite common; however, there are other solutions that are not quite as obvious. Solutions of metals, like brass, are called **alloys**. Alloys are formed by mixing molten metals and allowing them to solidify. Metals are mixed to take advantage of their individual properties like malleability, density, strength, and resistance to oxidation.

For example, steel is a mixture of iron with about 1% carbon. Pure iron, which is quite brittle, is mixed with carbon to give it greater strength. Stainless steel has metals like chromium and manganese added to the steel to decrease the rate of rusting. Metals such as copper are added to gold for jewellery so that the gold is not so soft. Aluminum alloys are used in bikes and other equipment that require strength, but must be lightweight as well.

Another type of solution mentioned in the table above is the amalgam. An **amalgam**, such as those used in silver dental fillings, is a mixture of mercury and an alloy of tin, silver, and copper. The mercury binds these metals together in a way that enables them to be manipulated into the cavity in a tooth.

All gases mix to form homogeneous mixtures, so all mixtures of gases are solutions. Gases such as hydrogen will also dissolve in some metals. Palladium metal, a solid, is useful in purifying hydrogen gas, since palladium can dissolve as much as 900 times its volume in hydrogen.

Mixtures—An Overview

The table below summarizes the differences between heterogeneous and homogeneous mixtures:

Type of Mixture	Homogeneous	Heterogeneous	
Property	**Solution**	**Colloid**	**Suspension**
Particle Type	Ions, atoms, small molecules	Large molecules or particles	Large particles or aggregates
Particle Size	Very small (0.1–1 nm)	Intermediate (1–1000 nm)	Large (greater than 1000 nm)
Stability	Stable (will not settle out)	Unstable (can be emulsified)	Unstable (can be emulsified)
Filtration	Solute cannot be filtered out	Particles cannot be easily filtered out	Filter paper can capture these particles
Appearance	Transparent, may have colour, does not show the Tyndall Effect	Cloudy or milky, opaque	Cloudy, translucent sediment will form

Note: The nm (nanometer) is 10^{-9} m.

Learning Activity 5.1

What Is a Solution?

1. What are the properties of a solution?
2. Define what is meant by a solution.
3. Compare (how are they the same) and contrast (how are they different) solutions, suspensions, and colloids.
4. Distinguish between homogeneous and heterogeneous mixtures.

Lesson Summary

In this lesson, you learned that a solution is a homogeneous mixture of two or more substances that exist together in one phase. You now know that a solution is transparent, cannot be filtered, and will not settle out if left standing. There are nine types of solutions that can be classified based upon the phase of the solute and solvent. In the next lesson, you will continue to explore solutions, focusing specifically on water and its unique properties.

NOTES

Assignment 5.1

Properties of Solutions *(5 marks)*

1. Identify the following as being homogeneous or heterogeneous. *(2 marks)*
   a) Vegetable soup

   ____

   b) Coffee with sugar

   ____

2. For each of the following solution types, give one example OTHER than those provided in the lesson. *(3 marks)*
   a) Liquid in liquid

   ____

   b) Solid in liquid

   ____

   c) Gas in gas

   ____

NOTES

Lesson 2: The Structure of Water (1 hour)

Lesson Focus

SLO C11-4-02: Describe the structure of water in terms of electronegativity and the polarity of its chemical bonds.

Lesson Introduction

In the last lesson, you learned that water dissolves many of the substances with which it comes in contact. This ability to be such a universal solvent is just one of the reasons why water is such a unique substance. In this lesson, you will learn about the structure of water and how this structure determines the polarity of this essential liquid.

Electronegativity

When two of the same atoms form a covalent bond, like in H_2, the electrons are shared equally between the two atoms. This, however, is not always the case. Different atoms have differing abilities to attract electrons when an atom is in a compound. This ability, called **electronegativity**, is the attraction that an atom has for the shared electrons in a covalent bond. Electronegativity values, which have no unit, are calculated based on factors like **ionization energy** (a measure of an atom's ability to lose electrons) and **electron affinity** (a measure of an atom's ability to gain electrons). Often, these values are displayed on the Periodic Table.

There are trends in electronegativity values that can be predicted using the Periodic Table of the Elements. *Electronegativity values generally decrease from top to bottom within a group or family and generally increase from left to right across a period.* These trends do not hold true for the transition metals, whose values do not necessarily follow these patterns. Additionally, since the Noble Gases are already stable and do not form many compounds, they are not associated with any electronegativity values.

Learning Activity 5.2

Electronegativity

Using the Periodic Table provided below, or Appendix G, answer the following questions regarding electronegativity.

1. Which element is the MOST electronegative?
2. Which element is the LEAST electronegative?
3. What is the electronegativity value for Lithium? Gallium? Lead?
4. Generally, non-metals have ___________ electronegativity values, while metals have _________ values.

Electronegativity Increases (across →)

Electronegativity Increases (up ↑)

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 2.20																	2 He —
3 Li 0.97	4 Be 1.47											5 B 2.01	6 C 2.50	7 N 3.07	8 O 3.50	9 F 4.10	10 Ne —
11 Na 1.01	12 Mg 1.23											13 Al 1.47	14 Si 1.74	15 P 2.06	16 S 2.44	17 Cl 2.83	18 Ar —
19 K 0.91	20 Ca 1.04	21 Sc 1.20	22 Ti 1.32	23 V 1.45	24 Cr 1.56	25 Mn 1.60	26 Fe 1.64	27 Co 1.70	28 Ni 1.75	29 Cu 1.74	30 Zn 1.66	31 Ga 1.82	32 Ge 2.02	33 As 2.20	34 Se 2.48	35 Br 2.74	36 Kr —
37 Rb 0.89	38 Sr 0.99	39 Y 1.11	40 Zr 1.22	41 Nb 1.23	42 Mo 1.30	43 Tc 1.36	44 Ru 1.42	45 Rh 1.45	46 Pd 1.35	47 Ag 1.42	48 Cd 1.46	49 In 1.49	50 Sn 1.72	51 Sb 1.82	52 Te 2.01	53 I 2.21	54 Xe —
55 Cs 0.86	56 Ba 0.97	57–71	72 Hf 1.23	73 Ta 1.33	74 W 1.40	75 Re 1.46	76 Os 1.52	77 Ir 1.55	78 Pt 1.44	79 Au 1.42	80 Hg 1.44	81 Tl 1.44	82 Pb 1.55	83 Bi 1.67	84 Po 1.76	85 At 1.96	86 Rn —
87 Fr 0.86	88 Ra 0.97	89–103	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —	110 Ds —	111 Rg —	112 Cn —	113 Uut —	114 Uuq —	115 Uup —	116 Uuh —		118 Uuo —

Inner Transition Elements

Lanthanide Series	57 La 1.08	58 Ce 1.08	59 Pr 1.07	60 Nd 1.07	61 Pm 1.07	62 Sm 1.07	63 Eu 1.01	64 Gd 1.11	65 Tb 1.10	66 Dy 1.10	67 Ho 1.10	68 Er 1.11	69 Tm 1.11	70 Yb 1.06	71 Lu 1.14
Actinide Series	89 Ac 1.00	90 Th 1.01	91 Pa 1.14	92 U 1.30	93 Np 1.29	94 Pu 1.25	95 Am —	96 Cm —	97 Bk —	98 Cf —	99 Es —	100 Fm —	101 Md —	102 No —	103 Lr —

Electronegativity and Bond Formation

When atoms with different electronegativities form a bond, the atom with the higher electronegativity value will exert a stronger pull on the shared electrons. If electrons are closer to one atom than another in a bond, the charge of the electrons is not shared equally. As a result, one atom “feels” more negatively charged than the other.

Bonds formed between atoms of differing electronegativities are called **polar bonds**, because one end of the bond has a partial positive charge and the other has a partial negative charge because the electrons are shared unequally. A good example of a polar bond is hydrogen fluoride, HF:

$$H \quad \delta+ \quad \longmapsto \quad \delta- \quad F$$

Fluorine is the more electronegative of the two atoms (4.10 vs. 2.20). When bonded to hydrogen, the fluorine draws the shared electrons towards itself. The arrow between the H and F atoms indicates the direction towards which the electrons are drawn, resulting in a more negative end. As a result, the fluorine end of the bond has a partial negative charge (δ–) and the hydrogen has a partial positive charge (δ+). These partial charges are electrically charged regions, or poles. This is why a molecule that has two poles is called a **dipolar molecule**, or **dipole**.

Since the H-F bond is polar, and this is a diatomic (two atoms) molecule, the molecule is also polar. Therefore, HF is a **polar molecule**. Since those atoms whose electronegativities differ by less than 0.2 form non-polar bonds, most diatomic molecules made of different atoms are polar molecules. Diatomic molecules made of the same atoms, such as H_2 and O_2, share electrons equally. These contain **nonpolar bonds** and are called **nonpolar molecules**. See the following summary table to review the characteristics of ionic and covalent bonds:

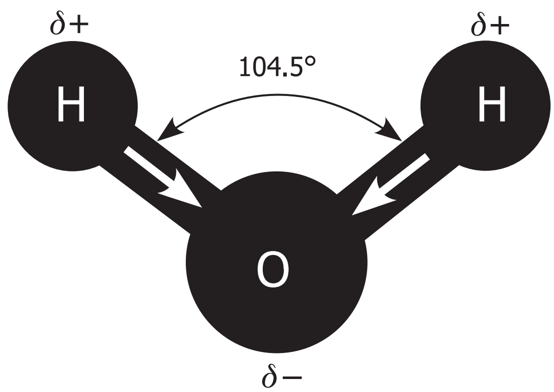
Type of Bond	Nature of Bond
Ionic	■ One or more electrons are transferred. ■ Ionic compounds are electrically neutral. ■ This type of bond forms between metals and non-metals. ■ Many ionic compounds are solid at room temperature and generally have high melting points. ■ Ionic compounds conduct an electric current when dissolved in water.
Covalent	■ Electron sharing occurs so that atoms attain a noble gas configuration. ■ Generally, this type of bond forms between non-metals. ■ Covalent molecules can be solid, liquid, or gas. ■ They are poor conductors and have low solubility in water and low melting points.
Non-polar Covalent	■ The atoms in the bond pull equally. ■ The bonding electrons are equally shared. ■ Includes diatomic molecules like H_2.
Polar Covalent	■ Electrons are shared unequally. ■ The more electronegative atom attracts electrons more strongly, and gains a slightly negative charge.

Generally, the greater the electronegativity difference, the more polar the bond. If the electronegativity difference is large enough, the bonds are considered to be ionic. This is why we use the general rule that metal and non-metal atoms make ionic bonds, since metals usually have low electronegativities and non-metals usually have high electronegativities.

The Structure of Water

Water is made up of two hydrogen atoms covalently bound to a single oxygen atom. That is, the hydrogen atoms and the oxygen atom each share a pair of electrons.

For water, when a hydrogen atom and an oxygen atom share electrons, they do not share them evenly. Oxygen is more electronegative than the hydrogen, and as a result, the electrons between the hydrogen and oxygen atoms in each bond lie more towards the oxygen than they do towards the hydrogen. The hydrogen atoms are bonded to the oxygen at an angle of 104.5°, which gives the water molecule its characteristic bent shape. The shape of the water molecule combined with the two polar bonds makes water a polar molecule.



As a result, the oxygen atom acquires a partial negative charge (δ–) and the less electronegative hydrogen atoms acquire partial positive charges (δ+). This means that the bonds formed between oxygen and hydrogen are highly polar. Since the water molecule is bent, the two polar O-H bonds cannot cancel each other out, and so the water molecule as a whole is also polar.

Each partial positive end of the water molecule will attract to the negative end of another molecule. This intermolecular interaction among water molecules results in the formation of **hydrogen bonds**. Hydrogen bonding is stronger than the normal attractive forces that occur between neighbouring molecules. The following scale illustrates the relative strength of H bonding:

1 10 100

On this scale, 1 would represent normal intermolecular attractive forces and 100 would represent the intermolecular attractive forces between covalently bonded compounds. Hydrogen bonding would lie at 10 on this scale.

Many unique properties of water, such as its high surface tension and low vapour pressure, result from the formation of these hydrogen bonds.

Learning Activity 5.3

The Structure of Water

Answer the following questions about the structure of water.

1. What part of the water molecule has a partial negative charge? A partial positive charge?
2. Which element in water has the higher electronegativity value?
3. Which part of the water molecule can form hydrogen bonds?

Lesson Summary

In this lesson, you learned that electronegativity describes how electrons are shared in a bond. You also learned that a polar bond results when atoms of different electronegativities are bonded to each other. Water is a polar molecule due to its shape and the polar O-H bonds. In the next lesson, you will continue your study of solutions by learning about the dissolving process.

Lesson 3: The Solution Process (2 hours)

Lesson Focus

SLO C11-4-03: Explain the solution process of simple ionic and covalent compounds, using visual, particulate representations, and chemical equations.
Include: crystal structure, dissociation, hydration

SLO C11-4-04: Explain heat of solution with reference to specific applications.
Examples: cold packs, hot packs....

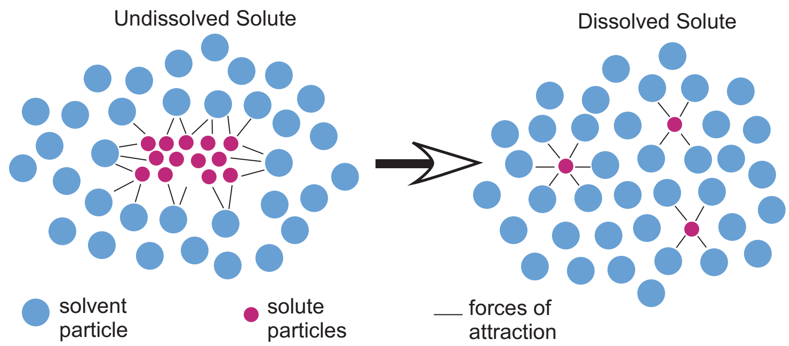
Lesson Introduction

In this lesson you will learn about the solution process—that is, how solutions are formed. As you will discover, not all compounds dissolve the same way. Finally, you will learn to identify dissolving processes that are exothermic and those that are endothermic.

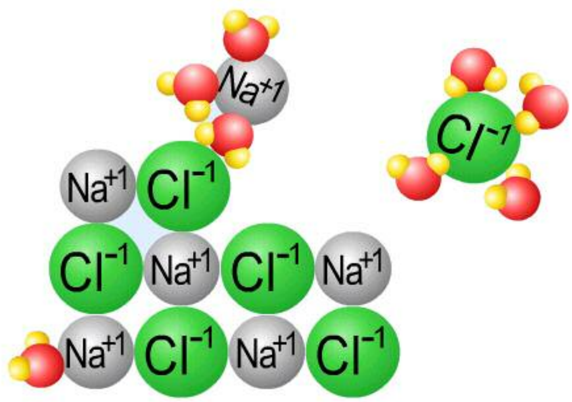
The Dissolving Process

How does the solution process, also known as the dissolving process, begin? Water molecules have kinetic energy, and so are in continuous motion. When a crystal of solute is placed in water, the water molecules collide with it from all sides. As the solute crystal dissolves in the solvent (water, in this case), the individual ions of the solute separate (for ionic compounds) or break away (for molecular compounds) from the surface of the solute crystals. The negatively and positively charged ions become surrounded by water molecules and the solute dissolves. Remember from the last lesson that a water molecule is polar, having a partial negative charge on the oxygen atom and partial negative charges on each hydrogen atom. Each of these partial positive charges on the hydrogen atoms will seek to form hydrogen bonds by attracting negatively charged solute ions.

The dissolving process depends on the solvent particles colliding with the solute ions, and on the forces of attraction between solute and solvent particles so that the solute particles are held in the spaces between solvent particles. See the diagram below showing the crystal structure of a solute as it dissolves.

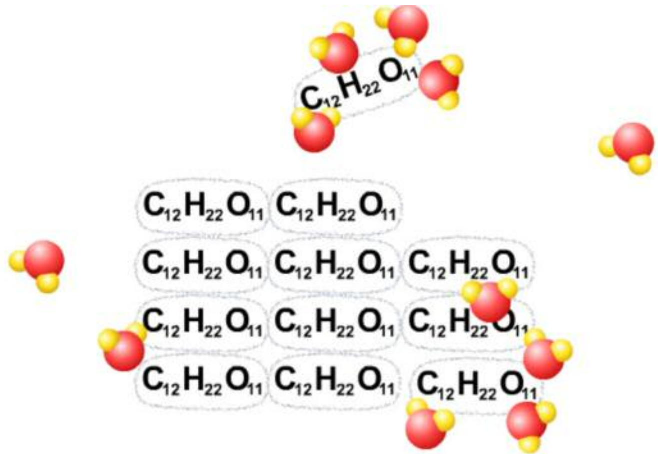


The illustration below shows how the dissolving of an *ionic compound*, like NaCl, would take place.



If you have access to the Internet, the dissolution process of an ionic compound is animated at <http://programs.northlandcollege.edu/biology/Biology1111/animations/dissolve.html>.

The following illustration shows how the dissolving of a *molecular compound*, like sugar, would take place.



What about solutes that do not dissolve when mixed with water? In some ionic compounds, the forces of attraction between the ions in the solute crystals are stronger than the attractions exerted by water molecules. In these cases, solvation does not occur and these compounds would be considered to be insoluble. Two such examples include barium sulfate ($BaSO_4$), used in x-ray imaging of the gastrointestinal tract, and calcium carbonate ($CaCO_3$), found in chalk and in antacids.

Steps in the Solution Process

There are three steps to the dissolving process:

1. *The solvent particles must separate to make room for solute particles.* This process requires energy to overcome the forces of attraction between solvent particles. The first step in the dissolving process is *endothermic*, as it requires an input of energy.

2. *The solute particles (or ions) must separate from the other solute particles at the surface of the crystals.* This process also requires energy to overcome the forces of attraction between the solute particles. Thus, the second step in the dissolving process is also *endothermic*.

   The energy of this step is known as **lattice energy**. Lattice energy is the amount of energy required to separate the molecules or ions from each other in a solid crystal.

3. When the solute ions move between the solvent particles *the forces of attraction between solute and solvent take hold and the particles "snap" back and move closer*. The final step in the dissolving process is *exothermic*, since this process releases energy.

The energy that is released in this final step is known as the **heat of hydration** or the **energy of hydration**. When water surrounds individual molecules or ions, the molecules or ions are said to be **hydrated**. The process by which positive and negative ions become surrounded by solvent molecules is called **solvation**.

Hydrates

You have surely heard the expression "stay hydrated," meaning make sure you are drinking enough fluids. In this module, hydration refers to the water contained in a crystal. A compound that contains water of hydration is called a hydrate. For example, copper (II) sulfate ($CuSO_4$) dissolved in water makes a deep blue, transparent solution. If you left this solution to evaporate, you would be left with deep blue crystals of copper (II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$). In writing the formula for a hydrate, a dot is used to connect the formula of the compound and the number of water molecules per formula unit.

Heat of Solution

The total heat change in the dissolving process is called the **heat of solution**. The heat of solution is equal to the sum of the heat changes for the three steps in the dissolving process. If the amount of energy absorbed in the first two steps of the dissolving process is *greater* than the amount of heat released, then the overall process is *endothermic*, an energy absorbing process. If the dissolving process for a substance is endothermic, the container will feel cooler as the substance dissolves. The dissolving of most solids in water, including ammonium nitrate, is an endothermic process:

$$NH_4NO_{3(s)} + \text{heat} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$$

If the amount of energy absorbed in the first two steps of the dissolving process is less than the amount of heat released, then the overall process is *exothermic*, an energy releasing process. The dissolution of calcium chloride is exothermic:

$$CaCl_{2(s)} \rightarrow Ca^{2+}{}_{(aq)} + 2Cl^-{}_{(aq)} + \text{heat}$$

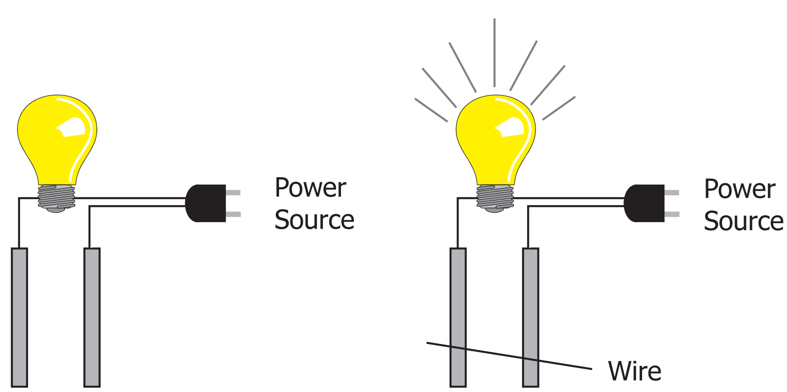
Heat of Solution Applications

Have you seen or used a cold pack that did not come from a freezer? The plastic pouch is at room temperature until you vigorously knead the bag in your hands, and then it begins to get cold! The endothermic dissolving process is used in these cold packs, which are often found in first aid kits. What is going on? The cold pack is a bag containing a solid, like ammonium nitrate, $NH_4NO_{3(s)}$, and a liquid, like water, that are initially sealed off, one from the other. The cold pack is activated by breaking the inner membrane and allowing the solid ammonium nitrate to mix with the liquid. The dissolving of ammonium nitrate absorbs large amounts of heat from its surroundings, including the skin against which it is placed.

Much like the cold pack, there are also hot packs that create heat by mixing two substances together. Remember that if the exothermic process is larger than the sum of each endothermic process, varying amounts of heat are released. Hot packs found in first aid kits will often contain a larger bag of solid calcium chloride, $CaCl_{2(s)}$, and a smaller inner bag of water. Breaking the inner bag releases water into the calcium chloride. As the calcium chloride dissolves, it releases heat. Other examples of solids used in hot packs are sodium hydroxide (NaOH) and potassium hydroxide (KOH).

Dissolving Sugar versus Salt in Water

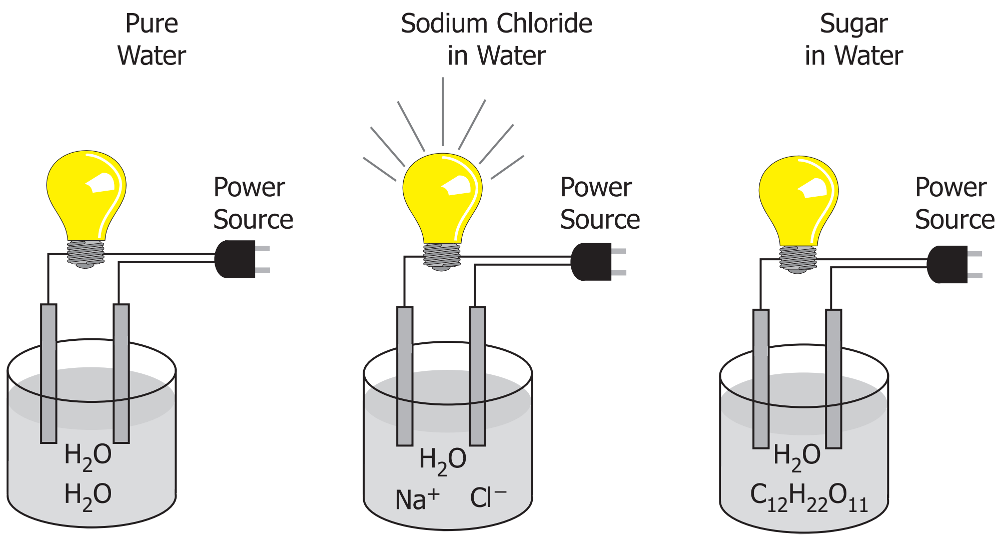
When you dissolve salt and sugar in water, there doesn't appear to be a visual difference between the two processes. However, if you look at the molecular level, a difference becomes apparent (see the diagrams of NaCl and sugar dissolving from earlier in this lesson). To observe a visual difference between the two types of solutions, look at the apparatus shown below.



The previous apparatus is simply a light bulb connected to two wire electrodes that interrupt an electrical current (electric charges—electrons or ions—in motion). The requirements for an electric current to flow are electric charges that are free to move, and a complete pathway around which these charges can travel. The light bulb does not light until the circuit is completed.

Note in the next diagram that if you placed the apparatus in a container of pure water, the light bulb would not light. This shows that pure water is NOT a conductor of electricity.

If the same apparatus were placed into a container with sodium chloride (table salt, NaCl) dissolved in water and then into a third container with sugar dissolved in water, it would be easy to see that only the NaCl solution conducts an electric current. If there were a fourth container with a more concentrated solution of salt, you would also observe a brighter light compared to the less concentrated solution of salt.



Only the sodium chloride solution conducts an electric current to light the light bulb. A solution of sodium chloride in water is called an **electrolyte**, since it conducts an electric current. However, a solution of sugar in water is a **non-electrolyte**, since it does not conduct an electric current. Pure water is a non-electrolyte because it does not contain anything that will carry an electric current. In order for a solution to conduct an electric current, charged particles or ions must be present in the solution and the charged particles must be free to move. The conductivity is due to the movement of the ions in solution between the two electrodes. Up to a certain limit, the more ions present in a solution, the greater the conductivity. Therefore, there are strong and weak electrolytes.

Sodium chloride is an ionic compound that breaks up into ions, or dissociates, when it dissolves. These ions become surrounded by water molecules (they are hydrated), and are free to move and carry an electric current through the solution.

What allows salt to conduct an electrical current, while sugar cannot? All ionic compounds (like NaCl) dissociate into positive and negative ions in water or, when melted, form liquids with charged particles that are free to move and conduct electric currents. Sugar, $C_{12}H_{22}O_{11}$, is a covalent solid that dissolves as whole molecules. These freely moving molecules are uncharged, so there are no particles that can carry a charge.

Dissociation Equations

The separation of positive and negative ions is called **dissociation**. Sodium chloride dissociates according to the equation shown below:

$$NaCl_{(s)} \xrightarrow{H_2O_{(l)}} Na^+_{(aq)} + Cl^-_{(aq)}$$

When ionic compounds like NaCl dissolve in water, they no longer exist as a compound, but as freely moving ions that are generally not associated with each other.

Molecular compounds like sugar do not separate into charged particles, but dissolve as whole molecules, and therefore do *not* dissociate. A non-electrolyte such as table sugar (sucrose, $C_{11}H_{22}O_{11}$) dissolves according to the following equation:

$$C_{11}H_{22}O_{11(s)} \rightarrow C_{11}H_{22}O_{11(aq)}$$

Example 1

Write the equation for dissolving solid magnesium chloride, $MgCl_2$, in water.

Step 1: *Determine if the compound is ionic or molecular.* Remember that molecular compounds do not dissociate.

Magnesium chloride is composed of metal (Mg) and non-metal (Cl) atoms. Since metal and non-metal atoms will usually produce an ionic compound, you can assume that there is dissociation.

Step 2: *Write the ions that will be formed in the dissociation.*

$$MgCl_{2(s)} \rightarrow Mg^{2+} + Cl^-$$

Step 3: *Use the subscripts from the formula to indicate the coefficients that will balance the charges.* (Remember what you did in Module 3?)

1 magnesium ion will be required to cancel the charge of 2 chloride ions. In the chemical equation, you do not need to write a coefficient if it is 1.

$$MgCl_{2(s)} \rightarrow 1Mg^{2+} + 2Cl^{-}$$

Step 4: *Write the equation using the appropriate state for the compound, (s) or (aq), for each dissociated ion.*

$$MgCl_{2(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

When magnesium chloride is dissolved in water, the compound dissociates into freely moving magnesium and chloride ions.

Example 2

Write the equation for solid aluminum sulfate, $Al_2(SO_4)_3$, dissolved in water.

Step 1: *Determine if the compound is ionic or molecular.*

Aluminum sulfate is composed of a metal, Al, and two non-metals, S and O. Since metal and non-metal atoms will usually produce an ionic compound, you can assume that there is dissociation.

Step 2: *Write the ions that will be formed in the dissociation.* A good starting point is to separate the metals from the non-metals, and then look up the corresponding charges.

$$Al_2(SO_4)_{3(s)} \rightarrow Al^{3+} + SO_4{}^{2-}$$

Step 3: *Use the subscripts from the formula to indicate the coefficients that will balance the charges.*

2 aluminum ions are required to cancel the charge of 3 sulfate ions.

$$Al_2(SO_4)_{3(s)} \rightarrow 2Al^{3+} + 3SO_4{}^{2-}$$

Step 4: *Write the equation using the appropriate state for the compound, (s) or (aq), for each dissociated ion.*

$$Al_2(SO_4)_{3(s)} \rightarrow 2Al^{3+}{}_{(aq)} + 3SO_4{}^{2-}{}_{(aq)}$$

Example 3

Write the equation for the dissolving of liquid methanol, CH_3OH, in water.

Step 1: *Determine if the compound is ionic or molecular.*

Methanol is composed entirely of non-metal atoms, which is usually an indication of a molecular compound. Molecular compounds do not dissociate, but dissolve as whole molecules.

> **Note:** The exception to this general rule is compounds containing the ammonium ion, NH_4^+. Compounds containing the ammonium ion, like ammonium chloride (NH_4Cl), are ionic compounds.

Step 2: *Write the equation using the appropriate state for the compound, (l) or (aq), for the dissolved molecule.*

$$CH_3OH_{(l)} \rightarrow CH_3OH_{(aq)}$$

Remember: Nonpolar compounds like NO and I_2 will not dissolve in water.

Learning Activity 5.4

Dissolving Compounds

1. What type of compound will dissociate in water? What type will not?
2. What part of the water molecule would be attracted to a solute ion such as Cl^-? Na^+?
3. How can you determine if the heat of reaction is endothermic or exothermic when looking at the chemical equation?
4. Write the equation for the dissolving of each of the following in water:
   a) $PbSO_{4(s)}$
   b) $C_6H_{12}O_{6(s)}$
   c) $KBr_{(s)}$
   d) $NaF_{(s)}$
   e) $CH_3OH_{(l)}$
   f) calcium chloride$_{(s)}$
   g) $Na_2CO_{3(s)}$
5. Describe what occurs during the process of an ionic solid dissolving into water. Include what happens to the solute and to the solvent.

Lesson Summary

In this lesson, you learned the three steps of the solution process. At the end of these three steps, you learned to differentiate an exothermic dissolving process and an endothermic dissolving process. Finally, you are now able to write the dissociation equations for ionic and molecular compounds. In the next lesson, you will take this knowledge of the solution process and perform a lab to illustrate the formation of solutions.

Assignment 5.2

The Solution Process *(10 marks)*

1. Write balanced dissociation equations for each of the following. Be sure to include states for the reactants and products. *(4 marks)*

   a) $C_{12}H_{22}O_{11(s)}$

   ____

   b) iron (III) sulfate$_{(s)}$

   ____

   c) potassium permanganate$_{(s)}$

   ____

   d) magnesium sulfide$_{(s)}$

   ____

2. How is the process of dissolving sugar different than the process of dissolving salt? Explain the reasons for these differences, basing your explanation on the nature of each compound. *(4 marks)*

   ____

   ____

   ____

   ____

   ____

   ____

continued

Assignment 5.2: The Solution Process (continued)

3. What **MUST** happen for an ionic compound to dissolve? *(2 marks)*

Lesson 4: Dissolving Polar and Non-Polar Substances (1.5 hours)

Lesson Focus

SLO C11-4-05: Perform a lab to illustrate the formation of solutions in terms of the polar and non-polar nature of substances.
Include: soluble, insoluble, miscible, immiscible

Lesson Introduction

In the last lesson, you learned that water is a polar molecule due to both its shape and the electronegativity difference between hydrogen and oxygen. In this activity, you will perform a lab to illustrate the formation of polar and non-polar solutions.

Lab Activity: Dissolving Polar and Non-Polar Substances

In Lesson 1 of this module, you learned that if a substance like sugar is able to dissolve in a solvent like water, the substance is said to be soluble. Likewise, if a substance like oil does not dissolve in a solvent like water, then the substance is said to be insoluble. Remember that very small amounts of an insoluble substance can often dissolve in a solvent, but since the amount is not significant, these substances are still considered to be insoluble. Finally, you learned that when two liquids dissolve in each other in any proportion, like alcohol and water, the liquids are miscible. Liquids that are insoluble in one another (such as oil and water) are immiscible.

Next, in Lesson 2, you learned about polar and non-polar substances. Recall that bonds formed between atoms of differing electronegativities are called polar bonds, because an unequal sharing of electrons results in one end of the bond having a partial positive charge and the other end of the bond having a partial negative charge. The example given of a polar bond was hydrogen fluoride, HF. Since the H-F bond is polar, the HF is considered to be a polar molecule. You learned that many diatomic molecules made up of *different* atoms are polar molecules. However, diatomic molecules made up of the same atoms, such as H_2 and O_2, share electrons equally. These molecules contain non-polar bonds, and are therefore considered to be non-polar molecules. How does the polarity of a substance affect its solubility?

Polar and charged substances dissolve well in polar solvents because of the electrostatic attraction between opposite charges. The attraction between the positive and negative ends of the water molecules and the polar ends of the solute hold the polar solute in solution. Many ionic compounds, like NaCl, are very soluble in water because of the electrostatic interaction between water and the positive and negative ions in the compound. When sodium chloride (NaCl) dissolves in water, the positively charged sodium ions (Na^+) are attracted to the partially negative charges on the oxygen atoms in the water molecules and the negatively charged chloride ions (Cl^-) are attracted to the partially positive charges on the hydrogen atoms in water. You may have heard water referred to as "the universal solvent" because it can dissolve many substances.

Non-polar substances, like waxes and oils, are not soluble in water, because the forces of attraction between individual water molecules are stronger than the attraction between the water and the oil. Recall that energy is required to overcome the electrostatic forces of attraction between the polar water molecules. That is, hydrogen atoms with partial positive charges in one water molecule will be attracted to the partial negative charges of oxygen atoms in other water molecules. Since oil is not a polar substance, there is very little attraction between the water and oil molecules. The energy released when the oil mixes with the water is not sufficient to overcome the cohesion between water molecules.

However, oil is not insoluble in every solvent. For example, oil does dissolve well in kerosene or turpentine, because the forces of attraction between oil molecules are about as weak as the forces of attraction between kerosene molecules, another non-polar substance. It is not difficult to separate the molecules of kerosene to make room for oil molecules.

Styrofoam is another non-polar substance, and as such, is not soluble in water. This has posed an environmental problem, as discarded Styrofoam packing peanuts take a long time to break down and can end up releasing toxic chemicals into the environment. Shipping companies have recently begun using polar cornstarch packing peanuts, as they are completely biodegradable. If you have access to the Internet, Steve Spangler demonstrates why starch peanuts are a more eco-friendly packaging alternative at www.youtube.com/watch?v=PN2198hMboQ. You may also try doing a youtube search for *eco-friendly packing peanuts*.

Lab Activity: Solubility Rules

Now that you have reviewed some important concepts that pertain to this lesson, you can proceed with the lab. Note that you must perform the lab before you can complete the Assignment.

Purpose

In this lab, you will discover which substances, when mixed, will form solutions in order to create a general rule about solutions created from polar and non-polar substances.

Materials

If you do not have some of these materials, use what you have and modify the table accordingly. Contact your tutor/marker if you would like to discuss using other substances.

- Small transparent containers for mixing
- Spoons
- Polar substances: table salt (sodium chloride), water, vinegar
- Non-polar substances: vegetable oil, acetone, kerosene/turpentine

Procedure

1. Using the chart below as a guide, mix all possible combinations of the polar and non-polar substances.

   If the substance is a liquid, then use about 60 mL (2 tablespoons).

   If the solution is a solid to be mixed with a liquid, add about 1 mL (one-quarter of a teaspoon) of the solid to the container first, and then add the liquid.

2. Once the substances are added to a container, they should be swirled to mix.

   All mixtures, except those involving kerosene, acetone, or turpentine, can be disposed of down the drain. Read the instructions on the kerosene, acetone, or turpentine container for disposal guidelines.

CAUTION: Kerosene and turpentine are flammable. Keep them away from a flame and use them only in a ventilated area.

Learning Activity 5.5

Predicting Solubility

Write a paragraph outlining your predictions for this lab. Which combinations of substances do you think will form a solution? Which combinations will not? On what do you base these predictions?

Lesson Summary

From the lab activity in this lesson, you have hopefully learned that the general rule when dissolving substances is “like dissolves like.”

Assignment 5.3

Solubility of Polar and Non-Polar Substances *(18 marks)*

1. Complete the following Observation Table. Remember that the term *miscible* refers ***only*** to the property of liquids being able to mix in all proportions to form a homogeneous solution. *(10 x 1 mark = 10 marks)*

Solute	Solvent	Soluble or Insoluble	Miscible or Immiscible
Water	Vegetable oil		
Oil	Kerosene, turpentine, or acetone		
Kerosene, turpentine, or acetone	Salt (NaCl)		
Salt (NaCl)	Vegetable oil		
Kerosene, turpentine, or acetone	Water		
Salt (NaCl)	Water		
Salt (NaCl)	Vinegar		
Vinegar	Water		
Kerosene, turpentine, or acetone	Vinegar		
Vegetable oil	Vinegar		

continued

Assignment 5.3: Solubility of Polar and Non-Polar Substances (continued)

2. What did you observe regarding the mixing of polar and non-polar substances? *(2 marks)*

3. Based on the general rule, use the following statements to determine whether the following substances are polar or non-polar. *(1 mark)* State your reasoning for each combination. *(1 mark)*

   a) Cobalt (II) chloride crystals are soluble in water. *(2 marks)*

   b) Ammonia gas is soluble in water. *(2 marks)*

   c) TTE (trichlorotrifluoroethane) liquid is not miscible in water. *(2 marks)*

Lesson 5: The Solubility Curve (1.5 hours)

Lesson Focus

SLO C11-4-06: Construct, from experimental data, a solubility curve of a pure substance in water.

SLO C11-4-07: Differentiate among saturated, unsaturated, and supersaturated solutions.

Lesson Introduction

You have already learned that when two substances form a solution they are soluble, and, in the case of liquids, miscible. In this lesson, you will learn other ways to describe a solution, based on its saturation (the ability of the solute to dissolve into a given amount of solvent at a particular temperature). Next, you will apply these terms to a solubility curve that you will learn how to construct.

Saturated, Unsaturated, and Supersaturated Solutions

The precise amount of solute that a solvent holds at a given temperature is fixed, or constant. Using this as a comparison point, you can describe a solution in terms of how much solute *is* dissolved as compared to how much *can* dissolve. A solution that contains the maximum amount of solute for that amount of solvent at a particular temperature is said to be **saturated**. A saturated solution is generally very stable. If you add more solute to the solution, it will not dissolve. Rather, the extra solute will fall to the bottom of the container.

If a solution can hold more solute at that temperature, the solution is **unsaturated**. If additional solute is added to the solvent, it will dissolve until the solution is saturated. For example, if you add more sugar to an unsaturated sugar solution, the sugar will continue to dissolve until the solution becomes saturated.

A **supersaturated** solution contains more solute than the solvent can normally hold at that temperature. A supersaturated solution is very unstable. If you add just one crystal of solid to a supersaturated solution, much more solute **precipitates** out of the solution. To precipitate means to return to the undissolved state from the solution. A supersaturated solution can precipitate/crystallize quickly if it is disturbed.

Let’s draw an analogy with these saturation terms and the booking capacity of a hotel. Consider the number of rooms in the hotel to be like the solvent, and the people using the rooms to be like the solute. If the hotel still has empty rooms available and there is room for more people, you could consider this to be like an unsaturated solution, where more solute could be added. If all the rooms are booked to capacity, then this would be like a saturated solution, where no more solute could be added. However, if the hotel is overbooked and people are waiting in the lobby, then this would be like a supersaturated solution that is holding more solute than it should.

The Solubility Curve

A few lessons down the road, you will learn more about the effect that temperature has on the solubility of a substance. For the moment, it is sufficient to say that solubility varies with temperature. You can find the solubility of a substance at various temperatures and plot these data on a graph. The result is what is known as a **solubility curve**.

In this lesson, you will use data provided to you to construct a solubility curve. Each point on the curve you will construct represents a *saturated* solution of potassium nitrate (KNO_3) in water at a different temperature. The area *below* the curve represents quantities that produce an *unsaturated* solution at that temperature. The points *above* the curve tend to indicate a *supersaturated* solution (although they can indicate a saturated solution with some remaining, or undissolved, solute). Use the blank graph provided on the next page to construct your solubility curve.

This page is blank to allow for easier use of the graph paper on page 41.

Learning Activity 5.6

Plotting a Solubility Curve

Step 1: Plot solubility on the *y*-axis (or vertical) using the units of grams solute/100g H_20.

Step 2: Plot temperature in °C on the *x*-axis (or horizontal).

Step 3: Join the points together to create a curve. Note that you are NOT drawing a line of best fit.

Step 4: Title the graph "Solubility Curve for KNO_3".

Step 5: Indicate on the graph where the following terms might apply: saturated, unsaturated, and supersaturated.

Temperature (°C)	Solubility (g solute/100 g H_2O)
0	12
20	34
40	68
60	112
70	140

Lesson Summary

In this lesson, you learned how to construct a solubility curve. You also learned how to use the curve to interpret the saturation of potassium nitrate at any given temperature. In the next lesson, you will continue to use solubility graphs to solve problems.

NOTES

Assignment 5.4

Interpreting a Solubility Curve *(10 marks)*

Use the Solubility Curve for potassium nitrate to answer the following questions:

1. Estimate the solubility of KNO_3 at 25ºC and at 63ºC. *(2 marks)*

   ____

   ____

2. Estimate the temperature at which the solubility of potassium nitrate is 20 g/100 g and 150 g/100 g. *(2 marks)*

   ____

   ____

3. For the following data sets, indicate whether the solution is unsaturated, saturated, or supersaturated. *(4 marks)*

Temperature (°C)	**Solubility (g solute/100 g H_2O)**	**Saturation**
0	10	
20	50	
30	48	
60	60	

4. What happens to the solubility of KNO_3 as the temperature increases? Use examples from your graph to support your explanation. *(2 marks)*

   ____

   ____

NOTES

Lesson 6: Solving Solubility Problems
(2 hours)

Lesson Focus

SLO C11-4-08: Use a graph of solubility data to solve problems.

Lesson Introduction

In the last lesson, you learned how to construct and interpret a solubility curve. In this lesson, you will use solubility graphs to help you solve solubility problems.

Solubility Units

Solubility is the maximum amount of solute that can dissolve in a certain amount of solvent. It can be further defined as the amount of solute needed to make a saturated solution under certain conditions of temperature and pressure. The units of solubility are usually provided in terms of the mass (in grams) of solute, per 100 grams of solvent, as in the last lesson. Other possible units used for solubility are g/L (grams of solute per litre of solvent), g/100 mL (grams of solute per 100 millilitres of solvent), and mol/kg (moles of solute per one kilogram of solvent).

The solubility of a solute must be determined experimentally. One method used to determine the solubility of a substance at a specific temperature is to add a known mass of solute to a known mass (or volume) of solvent, and heat the mixture until all of the solute dissolves. When the solute has dissolved, the solution is cooled until the first signs of solid crystals appear. The temperature at this point is recorded and the solubility is converted to the desired units for solubility.

In this course, you will be using solubility graphs and theoretical data to determine the solubility of a solution and solve related problems. In the lab, you might add 36.0 g of salt (NaCl) to 100 g of water at room temperature (25°C), and find that all the salt dissolves. You then might add another 1.0 g more of NaCl to the water, finding that only 0.2 g of it dissolves. Using this information, you can determine that the precise solubility of NaCl at room temperature must be 36.2 g/100 g of water. In other words, the solubility of a substance is the amount of solute that dissolves in a given quantity of solvent at 25ºC. Here are some examples to get you started:

Example 1

If 25.0 g of a solute is the maximum amount of solute that can dissolve in 40.0 g of solvent at a certain temperature, what is the solubility in grams of solute/100 g of solvent?

For this type of problem, it is best to use the following formula:

$$\frac{\text{mass of solute}}{\text{mass or volume of solvent}} \times \text{mass or volume of solvent needed}$$

The first term gives the grams of solute per gram of solvent multiplied by 100, for 100 g of solvent.

To solve this problem,

$$\frac{25.0 \text{ g solute}}{40.0 \ \cancel{\text{g solvent}}} \times 100 \ \cancel{\text{g solvent}} = 62.5 \text{ g solute/100 g solvent}$$

You can also solve this problem by using ratios. The solubility of the solute remains constant in this question. This allows you to set up a ratio equating the two solubilities:

$$\frac{\text{mass of solute 1}}{\text{mass or volume of solvent 1}} = \frac{\text{mass of solute 2}}{\text{mass or volume of solvent 2}}$$

If you solve this problem using ratios, you would cross-multiply and isolate the variable (mass of solute).

$$\frac{25.0 \text{ g solute}}{40.0 \text{ g solvent}} = \frac{\text{mass of solute}}{100 \text{ g solvent}}$$

$$\frac{25.0 \text{ g solute}}{40.0 \ \cancel{\text{g solvent}}} \times 100 \ \cancel{\text{g solvent}} = 62.5 \text{ g solute/100 g of solvent}$$

You may choose the method that you like best to solve similar problems.

Example 2

If 30.1 g of a solute can dissolve in 350.0 g of water at a certain temperature, what is the solubility of the substance in g/100 g water?

$$\frac{30.1 \text{ g solute}}{350.0 \text{ } \cancel{\text{g solvent}}} \times 100 \text{ } \cancel{\text{g solvent}} = 8.60 \text{ g solute }/100 \text{ g water}$$

Or, you can set up the ratio and cross-multiply:

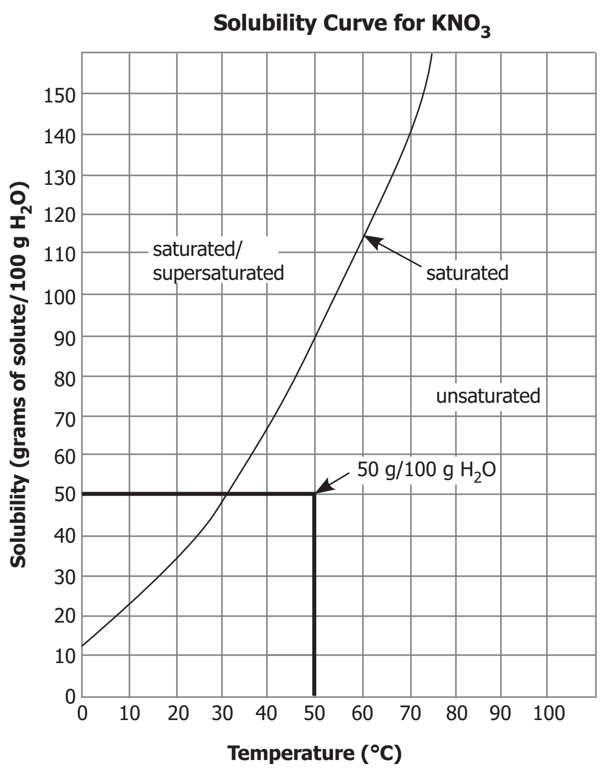
$$\frac{30.1 \text{ g solute}}{350.0 \text{ g solvent}} = \frac{\text{mass of solute}}{100 \text{ g solvent}}$$

$$\frac{30.1 \text{ g solute}}{350.0 \text{ } \cancel{\text{g solvent}}} \times 100 \text{ } \cancel{\text{g solvent}} = 8.60 \text{ g solute}/100 \text{ g of solvent}$$

Example 3

Using the solubility curve below, determine if a solution of 50 g of potassium nitrate in 100 g of water is saturated, unsaturated, or supersaturated at 50°C.

The *y*-axis indicates that the units of solubility are g/100 g water. Using the graph, is 50 g/100 g water a saturated solution? Start on the *x*-axis at 50°C and move up to the 50 g mark.

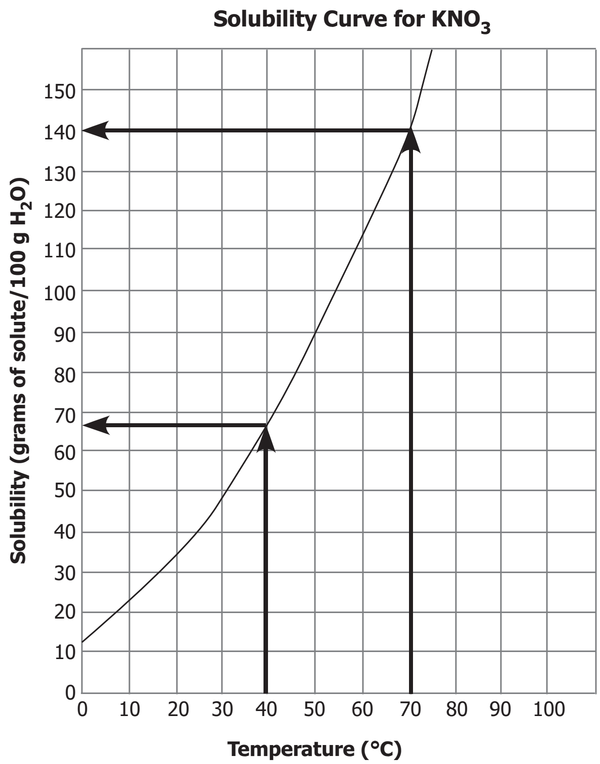


The solution is unsaturated, since 50 g of KNO_3 can dissolve in 100 g of water at 50°C.

Example 4

A 75 mL volume of a saturated solution of KNO_3 at 70°C is cooled to 40°C. Using the solubility curve for potassium nitrate, determine how much solid precipitates from the solution.

From the graph, we can see that a saturated solution of KNO_3 at 70°C contains 140 g/100 g water.



You only have 75 mL of water. Assuming the density of water is 1 g/mL, the mass of water is 75 g (for water, 1 g = 1 mL). Therefore, at 70°C:

$$\frac{140 \text{ g KNO}_3}{100 \cancel{\text{ g water}}} \times 75 \cancel{\text{ g water}} = 105 \text{ g KNO}_3/75 \text{ g water}$$

At 70°C, 105 g of solute will dissolve in 75 mL of water.

If the solution is cooled, the solubility of the solid will decrease. Reading the graph at 40°C, you can see that the solubility is now about 66 g/100 g water. We can convert this value to g/75 g water, using the same method as above:

$$\frac{66 \text{ g KNO}_3}{100 \cancel{\text{ g water}}} \times 75 \cancel{\text{ g water}} = 49.5 \approx 50. \text{ g KNO}_3/100 \text{ g water}$$

At 40°C, only 50. g of solute will dissolve in 75 mL of water.

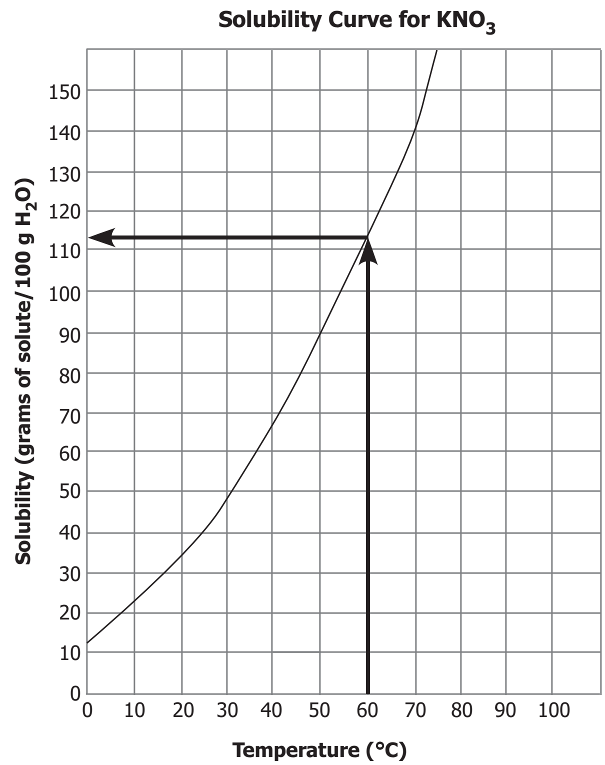
If only 50. g remain in solution and you started with 105 g, then the amount that comes out of solution is

$$105 \text{ g} - 50. \text{ g} = 55 \text{ g}$$

Therefore, 55 g will precipitate out of solution upon cooling.

Example 5

What volume of water is required to dissolve 240.0 g of potassium nitrate (KNO_3) at 60°C?



From the graph, at 60°C the solubility of KNO_3 is about 113 g/100 g water.

Using ratios,

$$\frac{113 \text{ g solute}}{100 \text{ g solvent}} = \frac{240.0 \text{ g solute}}{x \text{ g solvent}}$$

$$\frac{240.0 \cancel{\text{ g solute}}}{113 \cancel{\text{ g solute}}} \times 100 \text{ g solvent} = 212 \text{ g of solvent}$$

About 212 g of water (which is also 212 mL) of water is needed.

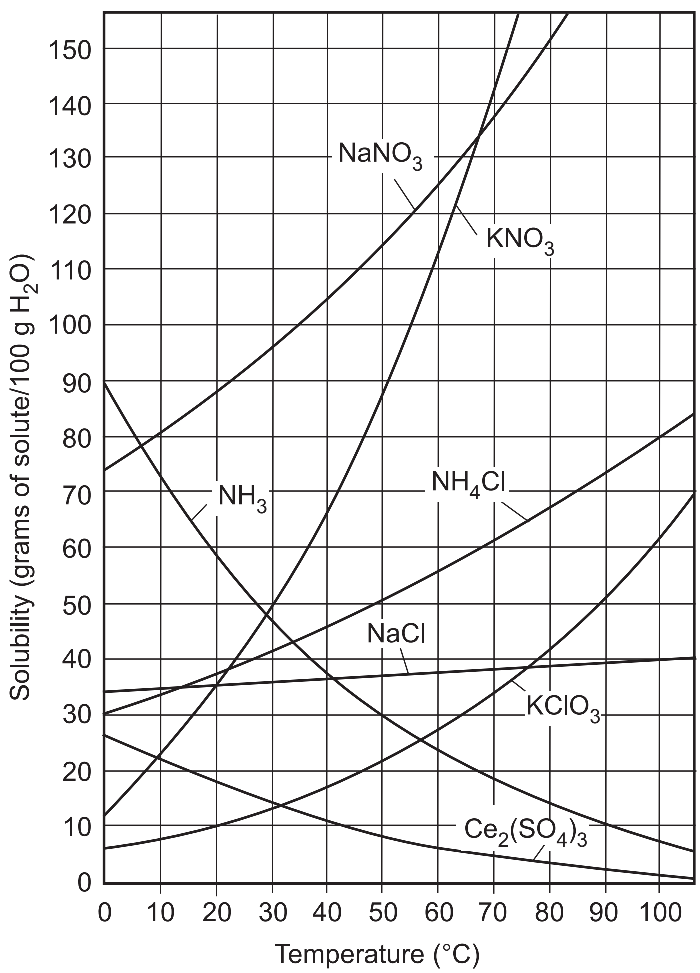
Remember, for water, you can convert from millilitres to grams using the conversion factor of 1 g = 1 mL.

Learning Activity 5.7

Solubility and Temperature

Use the Solubility versus Temperature for Several Substances graph below to help solve the following problems (except for Questions 1 and 2, which do not require the graph). Show your calculations where necessary.

Solubility versus Temperature for Several Substances



continued

Learning Activity 5.7 (continued)

1. Calculate the solubility of each of the following in g of solute/100 g of water.
   a) 0.250 kg dissolves in 1.2 L.
   b) 24.0 g dissolves in 280. g of water.
2. Determine the solubility of the following in g of solute/L water.
   a) 261 g of a solid dissolves in 1510 mL of water.
   b) 0.160 kg of a solid dissolves in 225 g of water.
3. Determine the temperature of the following substances using the given solubilities and the graph provided.
   a) KNO_3 120 g/100 g
   b) $NaNO_3$ 1200 g/L
4. What is the solubility, in g/100 g water, of the following solutions at the specified temperature?
   a) $NaNO_3$ at 40°C
   b) $Ce_2(SO_4)_3$ at 25°C
   c) NH_3 at 30°C
   d) NH_4Cl at 5°C
5. How much more NH_4Cl can you dissolve in a litre of water at 60°C than at 20°C?
6. If you prepared a saturated solution of $NaNO_3$ at 80°C and then cooled it to 30°C, what would happen? Be specific.
7. At which temperature do $NaNO_3$ and KNO_3 have the same solubility? What about NaCl and NH_3?
8. How much water is needed to dissolve 65.0 g of $NaNO_3$ at 35°C?
9. What temperature is necessary to dissolve twice as much KNO_3 as can be dissolved at 30°C?

continued

Learning Activity 5.7 (continued)

10. If the solubility of a solid in water is 118 g/L, how much water would you need to dissolve a piece of the same solid with a mass of 45.0 g?
11. If 18.0 g of KNO_3 are dissolved in 15.0 mL of water at 100°C, at what temperature will the solid begin to settle out?
12. If 40.0 g of KNO_3 is added to 50.0 mL of water at 40°C, will it all dissolve? If not, how much would be leftover? If you raised the temperature to 45°C, will it all dissolve? Give evidence.
13. If 142 g of NH_4Cl are dissolved in 350. mL of water at 55°C, is the solution saturated?

Lesson Summary

In this lesson, you used your knowledge of solution graphs to solve problems involving solubility and temperature. In the next lesson, you will learn more about how pressure and temperature affect the solubility of gases.

Assignment 5.5

Calculating Solubility *(14 marks)*

Use the Solubility versus Temperature for Several Substances graph from the previous Learning Activity to solve the following problems.

1. Calculate the solubility of each of the following in g of solute/100 g of water. (You do not need the graph for this question.) Show your calculations.
   a) 0.62 g dissolves in 15 mL of water. *(2 marks)*

   b) 75.0 g dissolves in 350. mL of water. *(2 marks)*

2. Determine the temperature of the following substances using the given solubilities and the graph provided in the lesson.
   a) NH_4Cl 60 g/100 g *(1 mark)*

   b) $KClO_3$ 100 g/500 g *(2 marks)*

continued

Assignment 5.5: Calculating Solubility (continued)

3. A saturated solution of $KClO_3$ in 200 g of water at 50°C is cooled to 20°C. Using data from the graph, determine how much $KClO_3$ will settle out. *(5 marks)*

4. Does solubility always increase with temperature? Give an example from the graph. *(2 marks)*

Lesson 7: Solubility of Gases (1 hour)

Lesson Focus

SLO C11-4-09: Explain how a change in temperature affects the solubility of gases.

SLO C11-4-10: Explain how a change in pressure affects the solubility of gases.

Lesson Introduction

How do you keep the "fizz" in your pop? Where's the best place to fish on a hot summer day? The answers to these questions can be answered by knowing about the solubility of gases in liquids. In this lesson, you will learn how temperature and pressure change the solubility of gases in liquid solvents.

Temperature and Solubility

In the last lesson, you learned that a higher temperature often increases the amount of solid that can dissolve in a liquid. This is due to the energy required during the dissolving process. Adding heat supplies more energy to separate the solute particles and solvent particles. Generally speaking, as more heat is added, more particles can enter into solution. However, the effect of temperature on the solubility of a gas in a liquid solvent is the opposite—there is a decrease in solubility as temperature increases. Thus, the solubility of most gases in liquid solvents is greater at colder temperatures than at warmer temperatures.

Why do the solubility rules change for gases? Increasing temperature decreases the solubility of a gas in a liquid because gases respond to temperature changes to a much larger extent than liquids. Remember that earlier in the course you learned gas particles are situated far apart from one another. An increased temperature increases the kinetic energy of the gas particles so that the liquid can no longer hold the gas particles in solution, and they "boil" out of the solution.

Temperature and Solubility Applications

You may have noticed that a carbonated beverage loses its fizz very easily when warmed. The fizz is caused by the carbon dioxide (the solute) dissolved in the soft drink returning to the gaseous state. Since the solubility of gases in water *decreases* with an increase in temperature, as the soft drink warms, the kinetic energy of the dissolved gas particles increases. As a result, the gas particles leave the solution, returning to the undissolved, or gaseous state. Thus, if you want to keep the "fizz" in your pop, keep your carbonated beverages cold.

Since the solubility of gases decreases as temperature increases, the heating of lakes and rivers during hot summers reduces the solubility of oxygen in the water. As fish and other aquatic animals rely upon dissolved oxygen to "breathe," the fish in small, shallow bodies of water may actually die during very hot summers due to a lack of oxygen. During hot summers, smart fishermen will drop their lines into the deepest (and thus coolest) portions of a lake where there will be the most oxygen. Like you and I, fish like to be where they can breathe more easily.

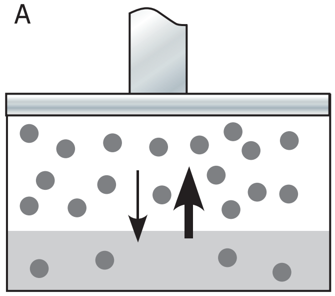
The same scenario occurs when an industrial plant draws water from a nearby lake, uses the water to cool its machinery, and then returns the water to the lake at a higher temperature. The temperature of the entire lake then increases, decreasing the oxygen solubility in the lake water. This phenomenon is known as **thermal pollution**.

Pressure and Solubility

Why is the solubility of carbon dioxide increased when a can of soda is under pressure? The general rule is that increasing pressure also increases the solubility of a gas in a liquid, because the increased pressure forces the gas into contact with the liquid. This added pressure forces large quantities of gas into solution and allows the liquid to hold onto the gas particles.

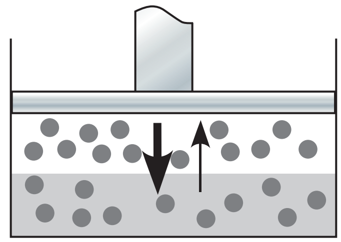
The pressure above a liquid affects the solubility of gases in that liquid. Carbon dioxide gas is bottled under 2 atm of pressure to keep it dissolved (see "B" in the following diagram). When you open a carbonated beverage (see "A" in the following diagram), the pressure is reduced to 1 atm. There is a sudden rush of gas from the can or bottle as the solubility of the carbon dioxide is reduced. Bubbles of gaseous carbon dioxide then form in the liquid.

A



1 atm of pressure

B



2 atm of pressure

Pressure and Solubility Applications

The effect of pressure on solubility is especially important to scuba divers. Underwater, for every 10 metres of depth, the pressure increases by 1 atmosphere. This means that the air a diver is breathing becomes more soluble and more readily absorbed by body tissues. Eventually, the pressure becomes so great that even the slightly soluble nitrogen gas found in air dissolves in a diver's blood. As the diver ascends, the decreasing pressure will cause a decrease in the solubility the dissolved gases in the blood. If the diver surfaces too quickly, the dissolved gases will begin to come out of solution and form bubbles in the blood vessels. This effect is compounded by the fact that these bubbles increase in volume as the pressure decreases (remember Boyle’s Law?). These bubbles can then become trapped in capillaries and block the flow of blood. This can be very painful or even deadly, especially if blood flowing to parts of the brain is blocked. This is what is known to scuba divers as the bends or decompression sickness. Therefore, it is suggested that divers ascend more slowly than the smallest bubbles escaping from their regulators, which equates to about 20 metres per minute. This allows excess nitrogen to be expelled by the respiratory system, rather than get trapped in the blood.

Learning Activity 5.8

The Effect of Pressure on Solubility

Complete the following Learning Activity to help you understand the effect of pressure on the solubility of a gas in a liquid. Use a bottle or can of your favourite soft drink to help you answer these questions.

1. Is this soft drink a solution? How do you know?
2. Look on the label. What are the solutes? What is the solvent? How do you know?
3. Are the solutes and solvent polar or non-polar?
4. Why does the drink make a sound when you open it?

Lesson Summary

In this lesson, you learned that the solubility of gases increases with decreasing temperature and increasing pressure. In the next lesson, you will perform a lab to demonstrate freezing-point depression and boiling-point elevation.

Lesson 8: Freezing-Point Depression and Boiling-Point Elevation (1.5 hours)

Lesson Focus

SLO C11-4-11: Perform a lab to demonstrate freezing-point depression and boiling-point elevation.

SLO C11-4-12: Explain freezing-point depression and boiling-point elevation at the molecular level.
Examples: antifreeze, road salt...

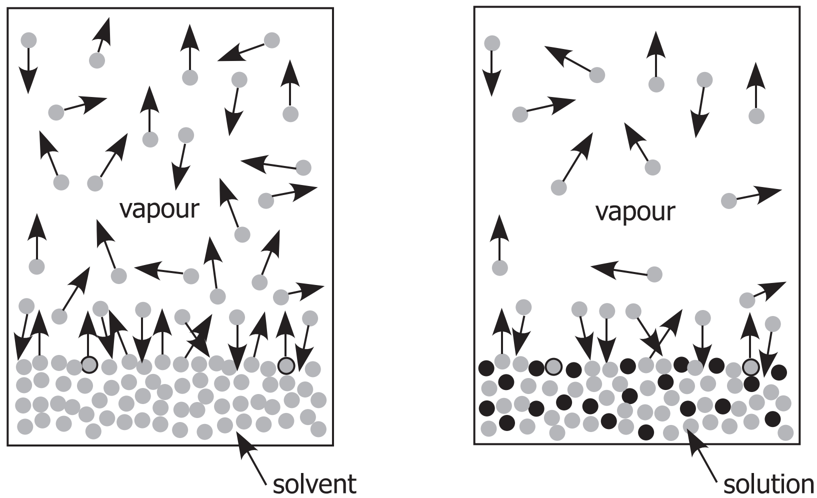
Lesson Introduction

You know that water boils at 100ºC and freezes/melts at 0ºC. In this lesson, you will learn that these temperatures can be slightly altered by changing the vapour pressure when a solute is added to water. You will also perform a short lab to prove that water doesn't always melt/freeze at zero degrees.

Effect of Solute on Vapour Pressure

Earlier in the course you learned that evaporation is a type of vaporization that occurs on the surface of a liquid. If a liquid is placed into a closed container, it will reach equilibrium between evaporation and condensation. At equilibrium, the pressure created by the vapour is called vapour pressure.

When a solute is added to a solvent, the solute becomes evenly distributed throughout the solvent as it dissolves. This means that some of the solute particles will occupy space at the surface of the liquid. As a result, the number of solvent particles at the surface is reduced.



Since the solvent particles can only evaporate at the surface of the liquid, and the solute prevents some of the solvent particles from reaching the surface, fewer solvent particles are able to evaporate. Fewer solvent particles at the solution's surface does not affect the rate of *condensation* of the vapour, though. The decrease in evaporation, along with the unchanged rate of condensation, results in a reduced amount of vapour and a lower vapour pressure. Therefore, *adding a solute to a solvent lowers the vapour pressure of the solvent*. This decrease in a solution's vapour pressure is proportional to the number of particles of solute in solution. In other words, the more solute that is added to a liquid, the greater the decrease in vapour pressure.

Freezing Point Depression

Up to this point, it has been sufficient to say that freezing point is the temperature at which solid and liquid are in equilibrium. The actual definition of freezing point (or melting point) is the temperature at which the vapour pressure of the solid and the vapour pressure of the liquid are equal. Since adding solute lowers the vapour pressure of the solvent, the temperature must be lowered for this equilibrium to occur in a solution. This is known as **freezing point depression**. The greater the amount of solute, the greater the freezing point depression (in other words, the lower the freezing point). For example, a 100 mL solution that contains 10 g of salt has a melting or freezing point of about −6°C, while the same volume with 20 g of salt can have a freezing point of −16°C.

Examples

Salt is spread on the streets in the winter to prevent icing. When cars drive on snow and ice, the pressure melts the ice and the salt dissolves in the water, lowering its melting point. This process of ice melting under pressure (which also occurs when you skate on ice or when you pack a snowball) is called **regelation**. The ice melts under pressure and then freezes when the pressure is released. This prevents the streets from icing up. Salt can only be spread on roads if the temperature is higher than −10°C, though, because it is not as effective at lower temperatures.

One of the salts of choice to prevent icing is calcium chloride, $CaCl_2$, because it dissolves exothermically, which melts even more ice. When calcium chloride dissolves, it forms 3 ions (Ca^{2+}, Cl^- and Cl^-), unlike the two ions formed by sodium chloride (Na^+ and Cl^-). Since freezing point depression is affected by the number of particles dissolved, calcium chloride is more effective as a de-icer on objects such as the wings of an airplane. It is also more environmentally friendly than sodium chloride.

Another example where solute reduces the freezing point of water is in sea water. Saltwater ports, like in Churchill, remain open even when the air temperature has been below zero for several days or weeks, partially due to the presence of salts in the ocean water.

Carbonated beverages are solutions that contain a number of different solutes (such as carbon dioxide and sugar) added to a water solvent. The presence of these solutes in water causes the freezing point of the pop to be lower than 0°C. This means that a sealed pop bottle can be cooled to temperatures below 0°C without it freezing. Opening the bottle of pop on a counter at room temperature will decrease the pressure on the carbon dioxide gas in the solution so that it can escape, lowering the number of solute particles, and increasing the freezing point so that the pop bottle freezes on the counter. This phenomenon is demonstrated online at www.metacafe.com/watch/382767/instantly_freeze_soda_experiment/.

Boiling Point Elevation

Earlier in the course, boiling point was defined as the temperature at which the vapour pressure of the liquid is equal to the air pressure. Since adding solute lowers the vapour pressure, more heat energy is then required to bring the vapour pressure up to the air pressure. This increases the temperature at which the solvent boils, resulting in **boiling point elevation**.

Boiling point elevation depends on the concentration of particles. Therefore, if more particles of solute are added, more kinetic energy is required to overcome the attractive forces that keep the solvent particles in the liquid and maintain the number of solvent particles evaporating. Also, since there are solute particles at the surface of the solution, it is less likely for the solvent particles to be able to escape from the solution. In other words, *the magnitude of the boiling point is proportional to the number of solute particles dissolved in the solvent*.

Examples

Sea water contains about 3.5 g of dissolved solids for every 100 mL of sea water. With this much solute, the boiling point increases only slightly, to about 100.6°C.

Radiator antifreeze is largely made of a compound called ethylene glycol. Ethylene glycol is mixed with water at a ratio of about 50:50 to increase the boiling point of the water in the radiator. The presence of the solute, along with circulating the coolant under pressure, increases the boiling point of water greatly. Thus, the coolant can absorb a lot of the engine's heat before it boils over. This is why you should never add just water if the coolant level in your car is low. Adding just water reduces the number of solute particles present, reducing the effectiveness of the coolant.

Colligative Properties

A property that depends on the number of particles of solute, and not upon their identity, is called a **colligative property**. There are three colligative properties that are important to understanding solutions: vapour pressure lowering, freezing point depression and boiling point elevation. In all three cases, the change is produced because of how much solute is added, regardless of the solute's identity.

Learning Activity 5.9

Colligative Properties

1. What factor determines how the vapour pressure, freezing point, and boiling point of a solution differ from those of a pure substance?
2. Would a dilute (weak) or a concentrated (strong) solution of saltwater have a higher boiling point? Why?
3. Equal quantities of KI and MgI_2 are dissolved into 1 L of water. Which solution will have the higher boiling point, and why?
4. What are the three colligative properties of solutions?

Lesson Summary

In this lesson, you learned that the amount of solute in a solvent can lower the vapour pressure, elevate the boiling point, and lower the freezing point of a solution. These are colligative properties. In the next lesson, you will learn about the concentration of a solution and how to calculate concentration.

NOTES

Assignment 5.6

Lab Activity: The Effect of Salt on the Melting of Ice *(10 marks)*

In this activity, you will investigate how the addition of salt affects the melting point of ice. Qualitatively, you could also observe this effect by simply freezing a volume of water and the same volume of an ice-salt (50/50) mixture. After the two samples are frozen solid, you could remove them from the freezer and watch which one melts faster.

Materials

Thermometer (You can ask a teacher to lend you one, or you can buy one at a hardware store.)
Coarse salt
Ice cubes—crushed
1 cup glass
Teaspoon

If you cannot obtain the above materials and you have access to the Internet, you may want to look up "freezing point depression animation" in any search engine. Effective online animations can be used to collect data for this experiment.

Procedure

1. Pour crushed ice into a 250 mL (1 cup) glass until it is half full.
2. Take the temperature of the ice while you stir the mixture with a spoon. Do NOT stir with the thermometer.
3. When the thermometer stabilizes (does not change any further), add one teaspoon of salt to the ice. Continue stirring. When the temperature stabilizes, record the temperature.
4. Repeat Step #3 until you have added 5 teaspoons, or until all of the ice has melted.

continued

Assignment 5.6: Lab Activity: The Effect of Salt on the Melting of Ice (continued)

Data Table

(0.5 marks x 6 = 3 marks)

Teaspoons of Salt Added	Temperature (°C)
0	
1	
2	
3	
4	
5	

Questions

1. What effect does salt have on the melting point of ice? Use your experimental data to support your explanation. *(2 marks)*

___

___

___

2. Explain why salt is added to ice-covered roads. *(1 mark)*

___

___

___

continued

Assignment 5.6: Lab Activity: The Effect of Salt on the Melting of Ice (continued)

3. Explain why antifreeze (ethylene glycol) is added to the water in the radiators of cars. Give an example of the freezing point of such a solution. *(2 marks)*

4. Which solution would have the higher boiling point—a solution of melted ice and salt, or the same volume of distilled water? Why? *(2 marks)*

NOTES

Lesson 9: Concentration (2 hours)

Lesson Focus

SLO C11-4-13: Differentiate among, and give examples of, the use of various representations of concentration.

SLO C11-4-14: Solve problems involving calculation for concentration, moles, mass, and volume.

Lesson Introduction

In recent years, the amount of liquid laundry detergent required to clean a load of clothes has become smaller. This is because the liquid laundry detergent is more concentrated. What does this mean? In this lesson, you will learn about the various representations of concentration and how to solve concentration problems.

Concentrated and Dilute

Where do you see the word concentration used? If you look at the label of many fruit juices, you will probably see the words "juice from concentrate." In the hospital, a patient receives an intravenous drip of a solution of salt or sugar mixed in water. The concentration of that solution is indicated on the bag. These are both examples of how a solution can be described in terms of how much solute and solvent are found in a specific solution.

If you have ever made a drink using powdered crystals, you have likely noticed that it is easy to make one glass very strong and sweet, while the next glass is weak and watery. **A concentrated solution** has a lot of solute as compared to the amount of solvent, while a **dilute solution** has little solute as compared to the amount of solvent present. In this case, the drink that is very sweet is the more concentrated solution of the two, and the drink that tastes watery is the more dilute solution.

Dilute and concentrated are relative terms. This can be illustrated using the three solutions below:

Solution A	Solution B	Solution C
500 g of sugar in 100 mL of water	100 g of sugar in 100 mL of water	1 g of sugar in 100 mL of water

Solution A is concentrated since it has a large amount of solute as compared to solvent. *Solution C* is dilute because it has a small amount of solute as compared to solvent. *Solution B* is dilute compared to solution A since it has less solute, but is concentrated compared to solution C because it has more solute than solution C.

Representations of Concentration

It is important for chemists to know how much solute is in a certain volume of solution. For example, chemists provide the data for safe concentrations of various pesticides, metals, and bacteria in drinking water. As you can imagine, if too much of these substances enters the drinking water, people may become sick. Therefore, testing drinking water to ensure safe concentrations of these contaminants is important for all living things in a community.

When chemists refer to **concentration**, they are referring to the amount of solute per unit of solution (solute plus solvent), unlike solubility that is the amount of solute per unit of solvent. There are several different units, or representations, for concentration. In this course, you will work with the following eight representations:

1. **Parts per million (ppm)** refers to the number of grams of solute for every one million grams of solution, or the number of particles of solute for every one million particles of solvent. For example, a solution that is 10 ppm of sodium ions in water means that there are 10 sodium ions for every one million water molecules.
2. **Parts per billion (ppb)** refers to the number of grams of solute for every one billion grams of solution, or the number of particles of solute for every one billion particles of solvent. For example, a water sample that is 10 ppb of iron ions in water means that there are 10 iron ions for every one billion water molecules.

The units of parts per million and billion are often used in describing concentrations of solutes present in small amounts in water and air. For example, after performing air quality tests, particles of pollutants in the air are often described in "parts per million."

Concentration can also be expressed using the following percentages:

3. **Percentage weight by weight (% w/w)** means the mass (in grams) of solute in 100 g of solution (weight/weight) or

$$\% \text{ w/w} = \frac{\text{grams of solute}}{\text{100 g solution}} \times 100\%$$

4. **Percentage weight by volume (%w/v)** means the mass (in grams) of solute for every 100 mL of solution (weight/volume) or

$$\% \text{ w/v} = \frac{\text{grams of solute}}{\text{100 mL solution}} \times 100\%$$

Both % w/w and %w/v are often used in medicine to describe the quantities of solutes, like cholesterol and iron, in the blood.

5. **Percentage volume by volume (%v/v)** means the volume (in mL) of solute for every 100 mL of solution (volume/volume) or

$$\% \text{ v/v} = \frac{\text{mL of solute}}{\text{100 mL solution}} \times 100\%$$

The percentage units are used in many industrial and household solutions. If you look at your household vinegar, you will see that it is 5% acetic acid.

Finally, concentration can also be represented using the following units:

6. Grams of solute in one litre of solution (g/L).
7. Grams of solute in every cubic decimeter of solution (g/dm^3).
8. Molarity: Read on to learn more about this representation of concentration.

Molarity

In chemistry, the most important unit of concentration is molarity, or molar concentration (M). The units most commonly used for molarity are:

- moles of solute in one litre of solution
- mol/L or,
- mol $\cdot$ L^{-1}

Note that while molarity is not an SI unit, like litres or grams, it is commonly used to represent concentration. For this reason, you need to know how to calculate molar concentration by dividing the moles of solute by the volume of the solution:

$$\text{Concentration } (C) = \frac{\text{moles of solute}}{\text{Litres of solution}}$$

$$C = \frac{n}{V}$$

Example 1

Calculate the concentration of a sodium chloride solution if 0.200 moles is dissolved in 250. mL of solution.

The units for concentration are moles per litre. You are given 0.200 moles of solute and 250. mL of solution. The volume of solution must be converted to litres.

$$\text{Volume (L)} = 250.\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.250\ \text{L}$$

Now substitute the given values into the equation for concentration:

$$\text{Concentration} = \frac{\text{moles}}{\text{Litres}} = \frac{0.200\ \text{mol}}{0.250\ \text{L}} = 0.800\ \text{mol/L}$$

0.200 moles of sodium chloride in 250. mL of solution has a concentration of 0.800 mol/L or 0.800 M (pronounced "molar").

Example 2

What volume of a 1.25 mol/L solution contains 5.00 moles of solute?

You are given a concentration of 1.25 mol/L and 5.00 moles of solute. To find the volume in litres, rearrange the equation to solve for volume and then substitute the given values into the equation.

$$\text{Concentration} = \frac{\text{moles}}{\text{Litres}}$$

$$\text{Litres} = \frac{\text{moles}}{\text{Concentration}} = \frac{5.00\ \cancel{\text{mol}}}{1.25\ \cancel{\text{mol}}/\text{L}} = 4.00\ \text{L}$$

4.00 L of a 1.25 mol/L solution contains 5.00 moles of solute.

Example 3

How many moles of solute are needed to make 400.0 mL of a 0.225 mol/L solution?

You are given the volume (in mL) and the concentration of the solution. First, convert the volume of the solution to litres, and then rearrange the equation to solve for moles so that you can substitute the given values into the equation.

$$\text{Volume (L)} = 400.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.4000\ \text{L}$$

$$\text{Concentration} = \frac{\text{moles}}{\text{Litres}}$$

$$\text{moles} = \text{Concentration} \times \text{Litres}$$

$$\frac{0.225\ \text{mol}}{1\ \cancel{\text{L}}} \times 0.4000\ \cancel{\text{L}} = 0.0900\ \text{mol}$$

0.0900 moles of solute are needed to make 400.0 mL of a 0.225 mol/L solution.

Calculating the Mass of Solute

When mixing a solution using a solid solute, you must first determine the mass of the solute needed to make a certain amount of solution.

Example 4

What mass of sodium chloride (NaCl) is needed to make 0.250 L of a 0.100 mol/L solution?

You are given:

- 0.250 L of solution (volume)
- 0.100 mol/L of solution (concentration)

Step 1: *Determine the number of moles of solute required.*

To find the number of moles, rearrange the concentration equation to solve for moles.

$$\text{moles} = \left(\frac{0.100\ \text{mol}}{\cancel{\text{L}}}\right)\left(0.250\ \cancel{\text{L}}\right) = 0.0250\ \text{moles of NaCl needed}$$

Step 2: *Convert from moles to mass after determining the molar mass of the solute.*

$NaCl$ = (1)(23.0 g/mol) + (1)(35.5 g/mol) = 58.5 g/mol

$$\text{mass} = \left(0.0250\ \cancel{\text{mole}}\right)\left(\frac{58.5\ \text{g}}{1\ \cancel{\text{mole}}}\right) = 1.46\ \text{g}$$

You need 1.46 g of sodium chloride to make 0.250 L of a 0.100 mol/L solution.

Example 5

What is the molar concentration of a solution of copper (II) sulfate ($CuSO_4$) if 50.0 g is dissolved in 600.0 mL of solution?

You are given:

- 50.0 g of solute (mass)
- 600.0 mL of solution (volume)

Step 1: *Convert from mass to moles after determining the molar mass of the solute.*

Since concentration is in moles per litre and not grams per litre, we must convert the given mass into moles using the molar mass of copper (II) sulfate.

$CuSO_4$ = (1)(63.5 g/mol) + (1)(32.1 g/mol) + (4)(16.0 g/mol)
= 159.6 g/mol

$$\text{moles} = 50.0\ \cancel{\text{g}} \times \left(\frac{1\ \text{mole}}{159.6\ \cancel{\text{g}}}\right) = 0.313\ \text{mol}$$

Step 2: *Determine the concentration of the solution.*

As the volume is given in mL, we must convert to L.

$$600.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.6000\ \text{L}$$

$$\text{Concentration} = \frac{\text{moles}}{\text{Litres}}$$

$$\text{Concentration} = \frac{0.313\ \text{mol}}{0.6000\ \text{L}} = 0.522\ \text{mol/L}$$

50.0 g of copper (II) sulfate dissolved in 600.0 mL of solution results in a concentration of 0.522 mol/L.

Other Sample Calculations

Example 1

Determine the concentration of a solution that contains 2.5 moles of solute dissolved in 5.0 L of solution.

$$\text{Concentration} = \frac{\text{moles}}{\text{L}} = \frac{2.5 \text{ moles}}{5.0 \text{ L}} = 0.50 \text{ mol/L}$$

Example 2

What volume of solution would be required to make a 0.40 mol/L solution dissolving 0.10 moles of solute?

$$\text{Volume (L)} = \frac{\text{moles}}{\text{concentration (mol/L)}}$$

$$\text{Volume} = \frac{0.10 \ \cancel{\text{mol}}}{0.40 \ \cancel{\text{mol}}/\text{L}}$$

The moles units cancel, leaving the answer in L.

Volume = 0.25 L (2 sig. figures)

Example 3

Calculate the mass of ammonium hydroxide, NH_4OH, required to prepare 0.30 L of a 0.25 mol/L solution.

$$\text{Moles} = \text{concentration} \times \text{litres}$$

$$= 0.25 \ \frac{\text{mol}}{\cancel{\text{L}}} \times 0.30 \ \cancel{\text{L}} = 0.075 \text{ moles}$$

$$\text{Mass} = 0.075 \ \cancel{\text{mol}} \times \frac{35.0 \text{ g}}{\cancel{\text{mol}}}$$

$$= 2.6 \text{ g (2 sig. figures)}$$

Example 4

Calculate the concentration of a solution if 44.0 g of lithium sulfate, Li_2SO_4, is dissolved in 0.400 L of solution.

$$\text{Moles} = \frac{44.0\ \cancel{\text{g}}}{109.8\ \cancel{\text{g}}/\text{mol}}$$

$$= 0.401\ \text{mol}$$

$$\text{Concentration} = \frac{\text{mol}}{\text{L}}$$

$$= \frac{0.40\ \text{mol}}{0.400\ \text{L}}$$

$$= 1.00\ \text{mol/L (3 sig. figures)}$$

Example 5

What volume would be required to make a 0.400 mole/L solution containing 51.01 g of sodium nitrate, $NaNO_{3(aq)}$?

$$\text{Moles} = \frac{51.01\ \cancel{\text{g}}}{85.0\ \cancel{\text{g}}/\text{mol}}$$

$$= 0.6001\ \text{mol}$$

$$\text{Litres} = \frac{\text{moles}}{\text{concentration}}$$

$$= \frac{0.6001\ \cancel{\text{mol}}}{0.400\ \cancel{\text{mol}}/\text{L}}$$

$$= 1.50\ \text{L (3 sig. figures)}$$

Learning Activity 5.10

Solving Concentration Problems

Solve the following problems, showing your calculations.

1. Calculate the concentration of a 200.0 mL solution that contains 0.250 moles of solute.
2. Find the concentration of a solution that contains 1.45 moles dissolved in 2.30 L of solution.
3. How many moles of NaOH would be needed to make 50.0 mL of a 0.750 mol/L solution?
4. What mass of $AgNO_3$ would be needed to make 250.0 mL of a 1.50 mol/L solution?
5. What mass of $CaCO_3$ would be needed to make 20.0 mL of a 0.400 mol/L solution?
6. How much solution would be needed to dissolve 50.0 g of $K2SO_4$ to make a 0.500 mol/L solution?
7. What volume of solution would be required to dissolve 18.04 g of aluminum sulfide to make a 0.160 mol/L solution?
8. What mass of sodium sulfate would be needed to make 50.0 mL of a 0.150 mol/L solution?

Lesson Summary

In this lesson, you learned that there are several ways to represent concentration, including the most commonly used representation, molarity. You then went on to solve problems involving calculations for concentration, moles, mass, and volume. In the next lesson, you will use these concepts to help you prepare a solution and determine its concentration.

NOTES

Assignment 5.7

Calculations Involving Concentrations *(11 marks)*

Solve the following concentration problems, showing your calculations. Remember to convert from mL to L when necessary.

1. A saline solution used to clean contact lenses contains 0.90 g NaCl in 100.0 mL of solution. Calculate the molarity of this solution. (Molar Mass of NaCl = 58.5 g/mol) *(3 marks)*

2. How many moles of ammonium nitrate (NH_4NO_3) are in 335 mL of a 0.425 mol/L solution of ammonium nitrate? *(3 marks)*

3. What mass of $CaCl_2$ would be used in 250 mL of solution to make a 2.0 mol/L solution of $CaCl_2$? (The molar mass of $CaCl_2$ is 111.1 g/mol) *(5 marks)*

Notes

LESSON 10: PREPARING A SOLUTION (1.5 HOURS)

Lesson Focus

SLO C11-4-15: Prepare a solution, given the amount of solute (in grams) and the volume of solution (in millilitres), and determine the concentration in moles/litre.

Lesson Introduction

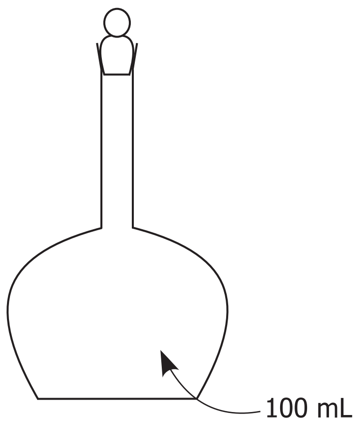
When preparing any solution, precision and accuracy is paramount. In this lesson, you will learn the process of preparing a solution, given the amount of solute (in grams) and the volume of solution (in millilitres). After preparing a solution, you will determine the concentration in moles of solute/litres of solvent.

Preparing a Solution

In the last lesson, you reflected on how you could prepare two glasses of drink that are not of the same concentration, based on the quantity of solute and solvent you added. It is important to note that you could end up with a more dilute glass of drink either by adding less solute (drink crystals) or by adding more solvent (water). Now imagine taking a full glass of water and adding several spoonfuls of sugar to make a solution. What would happen? If the glass is completely full, as you add sugar, the water would overflow. Because adding a solute will change the volume of the solute-solvent mixture, we cannot simply add 1 mole of solute to 1 L of solvent to make a 1 mol/L solution. The volume would no longer be 1 L.

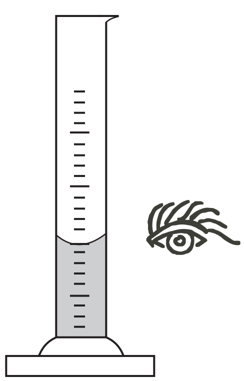
Here are the steps involved in making a solution with a solid solute and liquid solvent:

1. Obtain the precise mass of the solute needed.
2. Add the solid to a **volumetric flask** of the volume needed. Volumetric flasks come in various volumes, from as small as 10 mL to more than 2 L. If you want to make 100.0 mL of solution, you would choose a 100.0 mL volumetric flask.



3. Fill the **volumetric flask** about half full of solvent and swirl until the solid dissolves.
4. Carefully increase the level of solution until the correct level is reached, preferably with an eyedropper to ensure accuracy. It is most accurate to read the flask at eye level, rather than look down at the gradients.
5. Cover the top of the flask and shake the solution several times to mix the solute and solvent.

When measuring volumes with a graduated cylinder or volumetric flask, you must observe the **meniscus**. A meniscus is the curvature of the surface of the water. Water "sticks" to the walls of the graduated cylinder, but only on the sides and not the middle. If you looked at the liquid in a graduated cylinder, you would notice that the water level is not straight. Therefore, you should always read the measurement at the lowest point to get an accurate volume (see the diagram of the graduated cylinder below). In order to do this, you should always place the graduated cylinder or volumetric flask on a stable surface and then bend down to be able to read the meniscus at eye level.



If you have access to an accurate scale or balance, you can follow along with these steps to prepare the following solution.

Example 1

Describe the steps needed to make 500.0 mL of a 0.100 mol/L solution of sodium chloride (NaCl).

Step 1: *Determine the number of moles of solute needed.*

You are given volume in mL, so convert to L.

$$500.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.5000\ \text{L}$$

$$\text{moles} = \left(\frac{0.100\ \text{mol}}{\cancel{\text{L}}}\right)\left(0.5000\ \cancel{\text{L}}\right) = 0.0500\ \text{moles solute}$$

Step 2: *Calculate the mass after determining the molar mass of the solute.*

NaCl = (1)(23.0 g/mol) + (1)(35.5 g/mol) = 58.5 g/mol

$$\text{mass} = 0.0500\ \cancel{\text{mol}} \times \left(\frac{58.5\ \text{g}}{1\ \cancel{\text{mol}}}\right) = 2.92\ \text{g}$$

To summarize, the steps for making 500.0 mL of a 0.100 mol/L solution of NaCl are:

1. Mass out 2.92 g of NaCl.
2. Add the solid to a 500 mL volumetric flask.
3. Fill the flask half full with distilled water and swirl until the solid is dissolved.
4. Fill the flask to the 500 mL mark. Cap or seal the flask and swirl several times to mix.

To ensure the correct concentration of a solution, it is best to use distilled water, as there are fewer contaminants than in tap water. Finally, glassware that is used to prepare the solution should be clean.

If you have access to the Internet, a slightly altered version of the solution preparation process may be seen at www.youtube.com/watch?v=XMtm4hVCGWg.

Determining the Concentration

Once you have correctly prepared a solution, the next step would be to determine the concentration of that solution. In the last lesson, you learned that concentration is defined as the moles of solute divided by the volume of the solution:

$$\text{Concentration} = \frac{\text{moles of solute}}{\text{litres of solution}}$$

Remember to convert mL to L, and the mass of the solute to moles of the solute using the solute's molar mass.

Example 1

You prepare 100.0 mL of a solution using 5.85 g of sodium chloride. What would be the concentration of the prepared solution?

$$\text{Moles of NaCl} = 5.85\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{58.5\ \cancel{\text{g}}} = 0.100\ \text{mol NaCl}$$

$$\text{Concentration} = \frac{0.100\ \text{mol NaCl}}{0.1000\ \text{L}} = 1.00\ \text{mol/L}$$

Learning Activity 5.11

Preparing a Solution

1. When preparing a solution, why can you not obtain the required volume first and *then* add the mass of solute?
2. Why is it necessary to shake or swirl the solution well upon adding the final volume of solvent?
3. You prepare 250.0 mL of a solution using 11.5 g of potassium bromide, KBr. What is the concentration of the prepared solution?

Lesson Summary

In this lesson, you used your knowledge of concentration to calculate the concentration of a prepared solution. In the next lesson, you will continue learning about solutions, with a focus on solving problems involving dilutions.

NOTES

Assignment 5.8

Determining the Concentration of a Solution *(16 marks)*

1. Suggest *three* things you can do to ensure the most accurate preparation of a solution. *(3 marks)*

2. You prepare 400.0 mL of a solution using 24.3 g of potassium nitrate, KNO_3. What is the concentration of the prepared solution? (Molar mass of KNO_3 is 101.1 g/mol) *(4 marks)*

3. Describe the steps needed to make 250.0 mL of a 0.20 mol/L solution from solid NaOH. The molar mass of NaOH is 40.0 g/mol. A complete response should include the required mass of the sample, as well as a set of detailed instructions. *(9 marks)*

NOTES

LESSON 11: DILUTIONS (1.5 HOURS)

Lesson Focus

SLO C11-4-16: Solve problems involving the dilution of solutions.
Include: dilution of stock solutions, mixing common solution with different volumes and concentrations

SLO C11-4-17: Perform a dilution of a known concentration.

Lesson Introduction

When you take some frozen concentrated orange juice, put it into a pitcher and then add some water, you have performed a dilution. A can of condensed soup must also be diluted with water before you heat it on the stove. In this lesson, you will solve problems involving the dilution of solutions. You will also perform a dilution from a solution of known concentration.

Defining Dilution

Imagine that you have two 1.0 L solutions in beakers, and both solutions contain the same number of particles of solute. If one beaker then has more solvent added to it, it becomes the more *dilute* of the two solutions. You might even be able to make this conclusion by observation, as the more dilute solution may have a lighter colour than the more concentrated solution.

Chemists will often buy solutions from a chemical supply company in a more concentrated form so that they can save on storage space and more easily make new solutions. These types of more concentrated solutions are called **stock solutions**. Using a stock solution, you can make a solution of any lower concentration, simply by adding more solvent.

To dilute a solution means to add more solvent without adding more solute. *Diluting a solution reduces the number of moles of solute per unit volume, without changing the total number of moles of solute in solution.* Take the can of frozen orange juice as an example. If you add one can of water, you have diluted the concentrate. If you add four cans of water, the solution is much more diluted, even though the initial quantity of orange juice concentrate does not change.

If we add more solvent to a concentrated solution (let's call it solution 1), we will get a diluted solution (which we will call solution 2). The number of moles in both solutions remains unchanged, so the number of moles in solution 1, n_1, equals the number of moles in solution 2, n_2, or

$n_1 = n_2$.

We can also say:

Moles of solute before dilution = Moles of solute after dilution

Calculating Concentration

Example 1

If 500.0 mL of water is added to 300.0 mL of a 0.100 mol/L solution of NaCl, what is the new concentration?

Your givens are:

- The initial concentration, C_1, of the solution is 0.100 mol/L.
- The initial volume, V_1, of the solution is 300.0 mL.
- 500.0 mL of water was added to the solution.

Note: The added water is not the final volume. You must calculate the final volume after the dilution before you can use the value in your calculations.

Step 1: *Find the diluted volume.*

$$\begin{aligned} \text{Final volume } (V_2) &= \text{initial volume + added volume} \\ &= 300.0\text{ mL} + 500.0\text{ mL} = 800.0\text{ mL} \end{aligned}$$

The volumes should be converted to L:

$$\text{initial volume} = 300.0\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}} = 0.3000\text{ L}$$

$$\text{final volume} = 800.0\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}} = 0.8000\text{ L}$$

Step 2: *Use the dilution equation to solve for the final concentration.*

You already know from the previous lesson that the number of moles in a solution is given by

moles = Concentration x volume

or

$n = C \times V$

For the initial solution,

$$n_1 = C_1 \times V_1$$

where C_1 is the initial concentration in mol/L (stock solution), and V_1 is the initial volume of the solution in L (stock solution).

For the final solution,

$$n_2 = C_2 \times V_2$$

where C_2 is the concentration after dilution in mol/L (dilute solution), and V_2 is the volume after dilution in L (dilute solution).

We learned previously that the moles of solute before dilution are equal to the moles of solute after dilution, so

$$C_1 \times V_1 = C_2 \times V_2$$

or

$$C_1V_1 = C_2V_2$$

Note that volume can be in litres or millilitres, as long as the same unit is used for both V_1 and V_2.

So if

$$C_1V_1 = C_2V_2$$

$$(0.100 \text{ mol/L})(0.3000 \text{ L}) = C_2(0.80000 \text{ L})$$

$$\frac{(0.100 \text{ mol/L})(0.300 \cancel{\text{L}})}{0.800 \cancel{\text{L}}} = C_2$$

$$0.0375 \text{ mol/L} = C_2$$

The diluted concentration is 0.0375 mol/L.

Does the result make sense? Yes, as you are increasing the volume, you are diluting the concentration. The final concentration therefore should be less than your initial concentration.

Example 2

What is the concentration of the stock (initial) solution if 50.0 mL of the solution is diluted to make 220.0 mL of a 0.400 mol/L solution?

Your givens are:

- The final concentration, C_2, of the solution is 0.400 mol/L.
- The initial volume, V_1, of the solution is 50.0 mL.
- The final volume, V_2, of the solution is 220.0 mL.

Volume has increased, therefore diluting the concentration. You would expect the initial solution to be more concentrated than the diluted solution.

Step 1: *Find the diluted volume.*

$$50.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.0500\ \text{L}$$

$$220.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.2200\ \text{L}$$

Step 2: *Use the dilution equation to solve for the initial concentration.*

$$C_1V_1 = C_2V_2$$

$$C_1 = \frac{C_2V_2}{V_1} = \frac{(0.400\ \text{mol/L})(0.2200\ \cancel{\text{L}})}{0.0500\ \cancel{\text{L}}} = 1.76\ \text{mol/L}$$

The concentration of the stock solution is 1.76 mol/L.

Calculating Volume

Example 1

How much water must be added to 25.0 mL of a 1.00 mol/L stock solution of NaOH to make a 0.100 mol/L solution?

Your givens are:

- The initial concentration, C_1, of the solution is 1.00 mol/L.
- The initial volume, V_1, of the solution is 25.0 mL = 0.0250 L.
- The final concentration, C_2, of the solution is 0.100 mol/L.

Step 1: *Use the dilution equation to solve for the final volume.*

$C_1V_1 = C_2V_2$

$$V_2 = \frac{C_1V_1}{C_2} = \frac{(1.00\ \cancel{\text{mol/L}})(0.0250\ \text{L})}{0.100\ \cancel{\text{mol/L}}} = 0.250\ \text{L}$$

The *final* volume of the solution is 0.250 L, or 250. mL. This is *not* the volume of water added.

Step 2: *Calculate the volume of water added.*

Volume water added = Final Volume – Initial Volume
= 0.250 L – 0.0250 L = 0.225 L

The volume of water added is 0.225 L or 225 mL.

Example 2

How many millilitres of a 2.00 mol/L $MgSO_4$ solution must de diluted with water to prepare 100.0 mL of a 0.400 mol/L $MgSO_4$ solution?

Here are your givens:

- Initial concentration is 2.00 mol/L.
- Final Concentration is 0.400 mol/L.
- Final volume is exactly 100.0 mL.

You need to find the initial volume, which should be less than the final volume. *Why?* Because the initial solution is more concentrated than the final solution.

Step 1: *Use the dilution equation to solve for the initial volume.*

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{(0.400\ \cancel{\text{mol/L}})(100\ \text{mL})}{2.00\ \cancel{\text{mol/L}}} = 20.0\ \text{mL}$$

20.0 mL of the 2.00 mol/L $MgSO_4$ solution must be diluted with water to prepare 100.0 mL of a 0.400 mol/L $MgSO_4$ solution.

Learning Activity 5.12

Dilution Calculations, Part 1

Solve the following dilution problems, showing your calculation for each one.

1. Find the final concentration when 600.0 mL of 6.00 mol/L solution has 200.0 mL of water added to it.
2. A 0.500 mol/L solution has a volume of 8.00 L. The initial concentration was 2.50 mol/L. What was the initial volume?
3. You start with a solution that is 0.800 mol/L and 70.0 mL of water. You need to prepare a 0.300 mol/L solution. What is the final volume of the solution?
4. You have 500.0 mL of 3.00 mol/L sodium chloride solution. The solution is diluted to 580.0 mL. What is the final concentration?
5. 8.00 L of a 0.300 mol/L acid must be diluted to 0.0100 mol/L before it can safely be put into the sewage system. What is the final volume of the solution, and how much water must be added?
6. 60.0 mL of 6.00 mol/L sulfuric acid is diluted to 5.00 L. What is the final concentration?
7. Describe how to make 250.0 mL of a 0.100 mol/L solution of hydrochloric acid (HCl) from a solution that is 11.8 mol/L.
8. If 45.0 mL of stock HCl was required to make 150.0 mL of a 3.48 mol/L solution, calculate the concentration of the original stock solution.

Preparing a Dilution

Preparing a diluted solution is very similar to preparing a solution, except that you do not need to obtain the exact mass of a given solute.

To prepare a diluted solution:

1. Calculate the volume of stock solution needed. To do this, calculate the moles present in the one solution, and remember that the *moles in the more concentrated solution are equal to the moles in the more dilute solution*.
2. Add the stock solution to a volumetric flask of the appropriate size.
3. Add water to the volumetric flask up to the mark and mix.

If you have access to the Internet, you can observe the preparation of a diluted solution at www.youtube.com/watch?v=3EOiCrtvUUM.

Example 1

Describe how to make 1.00 L of a 1.00 mol/L solution of HCl from a concentrated solution of 11.8 mol/L.

You know:

- The initial concentration, C_1, of the solution is 11.8 mol/L.
- The final volume, V_2, of the solution is 1.00 L.
- The final concentration, C_2, of the solution is 1.00 mol/L.

Step 1: *Use the dilution equation to calculate the volume of stock solution needed.*

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{(1.00\ \cancel{\text{mol/L}})(1.00\ \text{L})}{11.8\ \cancel{\text{mol/L}}} = 0.0847\ \text{L}$$

You need 0.0847 L or 84.7 mL of the stock solution.

To make this solution, use a pipette to measure out 84.7 mL of the 11.8 mol/L HCl and add it to a 1 L volumetric flask. Fill the volumetric flask with distilled water to the mark and mix.

Example 2

How could you prepare 250.0 mL of a 0.20 mol/L solution of NaCl using a stock solution of 1.00 mol/L NaCl and water?

Step 1: *Use the dilution equation to calculate the volume of stock solution needed.*

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{\left(0.20\ \cancel{\text{mol/L}}\right)(250.0\ \text{mL})}{1.00\ \cancel{\text{mol/L}}} = 50.\ \text{mL}$$

To make this solution, use a pipette to measure out 50. mL of the 1.00 mol/L NaCl and add it to a 250 mL volumetric flask. Fill the volumetric flask with distilled water to the mark and mix.

Learning Activity 5.13

Preparing a Solution

1. How is preparing a dilution *different* from preparing a solution?
2. How is preparing a dilution *similar* to preparing a solution?

Mixing Solutions

Sometimes you may need to mix two or more solutions having the *same solute*, but *different concentrations* of that solute. If you are not sure whether the problem you are solving is one involving different concentrations, compare what characteristics are different before and after the solutions are mixed. Did you observe the following?

- Volumes are different ($V_1 \neq V_2$).
- Concentrations are different ($C_1 \neq C_2$).
- The solute is the same in both solutions.
- The solvent is the same in both solutions.

Since concentration is the number of moles of solute in one litre of solution, the final concentration of the mixture will be the total number of moles per litre of solution, or

$$C_{final} = \frac{\text{moles of solution 1} + \text{ moles of solution 2}}{V_1 + V_2}$$

or

$$C_{final} = \frac{C_1V_1 + C_2V_2}{V_1 + V_2}$$

where C_1 is the concentration of solution 1, in mol/L

V_1 is the volume of solution 1, in L

C_2 is the concentration of solution 2, in mol/L

V_2 is the volume of solution 2, in L

Example 1

Exactly 450. mL of a 0.150 mol/L NaCl solution is mixed with exactly 125 mL of a 0.220 mol/L NaCl solution. What is the new concentration of NaCl?

You know:

- The concentration of solution 1, C_1, is 0.150 mol/L
- The volume of solution 1, V_1, is 450. mL = 0.450 L
- The concentration of solution 2, C_2, is 0.220 mol/L
- The volume of solution 2, V_2, is 125 mL = 0.125 L

Add the total number of moles from both solution, and divide by the total volume:

$$\begin{aligned}
C_{final} &= \frac{C_1V_1 + C_2V_2}{V_1 + V_2} \\
&= \frac{(0.150 \text{ mol}/\cancel{\text{L}})(0.450 \ \cancel{\text{L}}) + (0.220 \text{ mol}/\cancel{\text{L}})(0.125 \ \cancel{\text{L}})}{0.450 \text{ L} + 0.125 \text{ L}} \\
&= \frac{0.0675 \text{ mol} + 0.0275 \text{ mol}}{0.575 \text{ L}} \\
&= \frac{0.0950 \text{ mol}}{0.575 \text{ L}} \\
&= 0.165 \text{ mol/L}
\end{aligned}$$

Notice that the final concentration is between the concentrations of the two original solutions. This is one quick check to make sure you have calculated correctly.

Example 2

What would be the final volume and concentration if 50.0 mL of a 0.250 mol/L NaOH solution is added to 75.0 mL of a 0.450 mol/L solution of NaOH?

The following procedure looks different from the solution to the previous example, but we are still finding the total number of moles and then dividing by the total volume. Below we have just calculated the number of moles as a separate step in the process.

$$\text{Moles of solution 1: } \frac{0.250\ \text{mol}}{\cancel{\text{L}}} \times 0.0500\ \cancel{\text{L}} = 0.0125\ \text{mol}$$

$$\text{Moles of solution 2: } \frac{0.450\ \text{mol}}{\cancel{\text{L}}} \times 0.0750\ \cancel{\text{L}} = 0.0338\ \text{mol}$$

$$\text{Final concentration} = \frac{\text{moles of solution 1} + \text{moles of solution 2}}{\text{volume 1} + \text{volume 2}}$$

$$= \frac{0.0125\ \text{mol} + 0.0338\ \text{mol}}{0.050\ \text{L} + 0.0750\ \text{L}} = 0.370\ \text{mol/L}$$

Learning Activity 5.14

Dilution Calculations, Part 2

1. What is the final concentration when a 1.0 L of a 0.10 mol/L solution is mixed with 1.0 L of a 1.0 mol/L of the same solution?
2. What is the final concentration when 400.0 mL of a 0.050 mol/L HCl solution is mixed with exactly 600.0 mL of a 0.020 mol/L HCl solution?

Lesson Summary

In this lesson. you learned that a diluted solution has the same number of moles of solute as the stock solution, but has a new volume. You practiced solving problems involving dilutions, including those where two or more solutions containing different concentrations of the same are mixed. You also reviewed the steps involved in preparing a dilution. In the next lesson, you will discover situations where solutions of known concentration are important.

NOTES

Assignment 5.9

Solving Dilution Problems *(15 marks)*

Solve the following dilution problems. For each problem, show the equation you are using and all calculations.

1. 450.0 mL of 2.40 mol/L H_2SO_4 is mixed with 375.0 mL of 8.20 mol/L H_2SO_4. What is the final concentration of H_2SO_4? *(4 marks)*

2. Find the final volume of a solution in which a 300.0 mL solution is diluted from 4.0 mol/L to 3.0 mol/L. *(3 marks)*

3. After dilution, a 1.70 mol/L solution has a volume 50.0 mL. If the original concentration of the solution was 2.00 mol/L, what was the solution's original volume? *(3 marks)*

continued

Assignment 5.9: Solving Dilution Problems (continued)

4. 47.5 mL of KI solution are needed to prepare 250.0 mL of a 0.760 mol/L solution of KI. What was the initial concentration of potassium iodide solution used? *(3 marks)*

5. Suggest two possible sources of error when preparing a dilution. *(2 marks)*

Lesson 12: Concentration Applications (1 hour)

Lesson Focus

SLO C11-4-18: Describe examples of situations where solutions of known concentration are important.
Examples: pharmaceutical preparations, administration of drugs, aquariums, swimming pool disinfectants, gas mixes for scuba, radiator antifreeze...

Lesson Introduction

Throughout this module, there have been references to concentrations and the importance of knowing the concentration of a solution. Some examples that were touched on include intravenous solutions for hospital use and concentrated beverages. In this lesson, you will learn of more situations where solutions of known concentration are important.

The Importance of Concentration Information

As you near the end of this module on solutions, you are no doubt realizing how precise measurement and using the correct concentration are important in many situations. As a patient receiving an intravenous solution of electrolytes, the wrong concentration can have adverse effects on your health. Likewise, if a lab technician incorrectly determines the concentration of components in your blood, you could receive the wrong type of care as a result. Read on to find out about more situations where knowing the concentration of a solution is particularly important.

Products Found at Home

Examples

- Vicks® VapoRub®: camphor 4.73% w/w
- Curél® moisturizer: glycerine 12% w/w
- rubbing alcohol: 40% v/v
- Aquafina® bottled water: fluoride ion, 0.3 ppm
- Roundup®: 7 g/L glyphosate
- liquid ant killer: 5.4% w/v borax

Pharmacy Preparations

Most ointments exceeding certain concentrations are dispensed by a qualified pharmacist.

Example

- Medicated dermatitis cream: 0.5% w/w betamethasone (an over-the-counter equivalent has a 0.1% w/w concentration)

Dental Surgery

Dentists will often use epinephrine in anaesthetic as a vasodilator to ensure that the anaesthetic is not flushed as rapidly from the tissues of the oral cavity.

Example

- Xylocaine hydrochloride 2% with 1:50,000 epinepherine (20 ppm)

Automobile Antifreeze

Example

- For one brand of commercial antifreeze/coolant, the container lists the following concentrations: For protection from freezing down to –52°C, dilute to a 60% v/v solution; and for further protection to –64°C, dilute to a 70% v/v solution.

Fish Aquariums

Nitrogen as ammonia in fish tanks must be carefully balanced to ensure fish remain healthy. The following table provides recommended concentrations in mg/L (ppm) of ammonia at various pH values.

pH	20°C	25°C
6.5	15.4	11.1
7.0	5.0	3.6
7.5	1.6	1.2
8.0	0.5	0.4
8.5	0.2	0.1

Reference: The Fishtank Crew: <http://pcpages.com/fish/nitrogen1.html>

Swimming Pool Solutions

The chlorine in swimming pools is carefully controlled as a disinfectant against bacteria and other micro-organisms.

Example

- Free chlorine, Cl_2, is usually kept between 1.0 and 2.5 ppm.

The use of chlorine as a disinfectant has come under scrutiny lately as a potential human biohazard. For background information, see the following website:

Centers for Disease Control and Prevention. *Facts about Chlorine.* www.bt.cdc.gov/agent/chlorine/basics/facts.asp.

Recreational Scuba

When a number of gases are placed together in the same container, the resulting system could be called either a gas mixture or a solution of miscible gases blended together. Scuba diving has become both safer and more complex with the use of Nitrox mixtures, in which the percentage of oxygen in air is increased and the percentage of nitrogen is decreased. By reducing the amount of nitrogen, divers are able to make longer dives to the same depth. Specialized instruction and certification is required for divers to use Nitrox mixtures. Excess nitrogen dissolved in the blood supply is the causal factor in *decompression sickness*, also known as the bends. The physiological effects of this situation often have outward manifestations such as severe muscle cramping and delirium, hence the reference to *bending*.

The following table, illustrating examples of gas mixtures, provides the maximum time permissible, for a particular depth, before decompression is required on ascent. EAN is Enriched Air Nitrox. EAN 32 is 32% O_2 as opposed to the normal 21% O_2 in air.

A Comparison of Dive Times for Various Gas Mixtures (Dive Times in Minutes)			
Depth (ft.)	**Air**	**EAN 32**	**EAN 36**
50	80	200	200
60	55	100	100
70	45	60	60
80	35	50	60
90	25	40	50
100	22	30	40
110	15	25	30
120	12	25	n/a

Reference: Oakley, Burks II. *Nitrox Scuba Diving.* University of Illinois at Springfield. 30 April 2000. http://www.online.uillinois.edu/oakley/nitrox.html.

Lesson Summary

In this lesson, you investigated some examples of solutions where the concentration was indicated. In the next lesson, you will focus on the treatment of our water supply.

Lesson 13: Our Water Supply (1.5 hours)

Lesson Focus

SLO C11-4-19: Describe the process of treating a water supply, identifying the allowable concentration of metallic and organic species in water suitable for consumption.

Lesson Introduction

Several times a day, you use drinking water without a second thought. The water that comes out of a tap has undergone a process to ensure that it is safe to drink and has allowable concentrations of contaminants. In this lesson, you will investigate the process of treating a water supply.

Water Treatment Process

Water treatment plants use a variety of treatment processes, including filtration, chlorination, and water softeners.

Filtration

Suspended solids and other fine particles can be removed from liquid streams by passing them through filters. This method is useful in removing sand, gravel, carbon, and other granular substances from water, but is not effective against bacteria and other microbes.

Chlorination

Chlorine dioxide (ClO_2) is used as a disinfectant in drinking water treatment. Chlorine dioxide is highly effective in controlling waterborne pathogens (disease-causing organisms) such as viruses, bacteria, and fungi, and many types of parasites. Chlorine dioxide is also an effective control strategy for taste, odour, colour, iron, and manganese removal.

Chlorine can be added to drinking water in one of several forms: elemental chlorine (chlorine gas), sodium hypochlorite solution (bleach), and dry calcium hypochlorite. Once added to water, each of these forms "free chlorine", which destroys pathogenic (disease-causing) organisms. Chlorine is also an effective control strategy for taste and odour. Chlorine disinfectants will also eliminate slime bacteria, moulds, and algae that commonly grow in water supply reservoirs, on the walls of watermains, and in storage tanks.

Use of Water Softeners

When water is in the ground, it picks up soluble particles of the material it passes through. Such soluble particles include magnesium and calcium, which are minerals commonly found in the Earth. When water contains high levels of these minerals, it is said to be hard water. Hard water does not pose a health risk, but can cause increased levels of soap scum and scale. Both of these are a nuisance to homeowners, as they can cause build-up and even blockage of water piping. Hard water does not lather well, and, as such, the action of cleaners, soaps, and shampoos in the home is reduced. A common water softener that is used is sodium chloride, NaCl, which ionizes and then forms compounds with the magnesium and the calcium. The calcium and magnesium ions in the water are replaced with sodium ions, forming calcium chloride ($CaCl_2$) and magnesium chloride ($MgCl_2$). Since sodium does not precipitate out in pipes or react badly with soap, both of the problems related to hard water are eliminated. Sometimes, a water softener is added directly to the water in the form of tablets or salts. In other instances, the water in the house runs through a bed of small plastic beads covered with sodium ions. As the water flows past the sodium ions, they swap places with the calcium and magnesium ions. The water with the sodium ions then travels on to the house.

Learning Activity 5.15

Water Treatment Methods

1. Research then summarize the following water treatment methods:
   a) Fluorination
   b) Distillation
   c) Aeration
   d) Reverse Osmosis
2. What is your local water source? See the Manitoba Water Stewardship Branch website at www.gov.mb.ca/waterstewardship.odw/ for information on your community water supply.
3. What does the term *potable water* mean? Give an example of non-potable water.

Lesson Summary

In this lesson, you researched some of the methods used to treat your water supply. You also learned about the water source(s) for your community.

NOTES

Module 5 Summary

Congratulations! You have reached the end of Module 5. You have only one module left.

Submitting Your Assignments

It is now time for you to submit Assignments 5.1 to 5.9 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 5 assignments and organize your material in the following order:

- Module 5 Cover Sheet (found at the end of the course Introduction)
- Assignment 5.1: Properties of Solutions
- Assignment 5.2: The Solution Process
- Assignment 5.3: Solubility of Polar and Non-Polar Substances
- Assignment 5.4: Interpreting a Solubility Curve
- Assignment 5.5: Calculating Solubility
- Assignment 5.6: Lab Activity: The Effect of Salt on the Melting of Ice
- Assignment 5.7: Calculations Involving Concentration
- Assignment 5.8: Determining the concentration of a Solution
- Assignment 5.9: Solving Dilution Problems

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

Notes

Grade 11 Chemistry (30S)

Module 5: Solutions

Learning Activity Answer Keys

Module 5: Solutions

Learning Activity 5.1: What Is a Solution?

1. What are the properties of a solution?

   Answer:

   A solution is homogeneous, exists in a single phase, its particles are too small to be seen, it is transparent, it does not settle out, and it cannot be separated by filtration.

2. Define what is meant by a solution.

   Answer:

   A solution is a homogeneous mixture of two or more substances, in one phase, where the substances are present as individual molecules or ions.

3. Compare (how are they the same) and contrast (how are they different) solutions, suspensions, and colloids.

   Answer:

Solutions	Suspensions	Colloids
Compare		
Mixture Cannot be filtered Does not settle	Mixture	Mixture Not easily filtered Does not settle
Contrast		
Homogeneous, not easily filtered, does not settle out, transparent, one phase visible, individual particles too small to be seen	Heterogeneous, easily filtered, settles out, opaque, more than one phase, particles visible	Heterogeneous, not easily filtered, does not settle out, translucent, more than one phase, particles may be visible

4. Distinguish between homogeneous and heterogeneous mixtures.

   Answer:

   While the particles in homogeneous mixtures are evenly distributed, those in heterogeneous mixtures are not evenly distributed. Individual parts of a heterogeneous mixture are visible, while the parts are not visible in a homogeneous mixture.

Learning Activity 5.2: Electronegativity

Using the Periodic Table provided below, or Appendix 5.1, answer the following questions regarding electronegativity.

1. Which element is the MOST electronegative?

   Answer:

   Fluorine is the most electronegative of all atoms with an electronegativity value of 4.10 (no units).

2. Which element is the LEAST electronegative?

   Answer:

   The least electronegative elements are Cesium and Francium, with electronegativity values of 0.86.

3. What is the electronegativity value for Lithium? Gallium? Lead?

   Answer:

   0.97, 1.82, 1.55

4. Generally, non-metals have ___________ electronegativity values, while metals have _________ values.

   Answer:

   Generally, non-metals have *high* electronegativity values, while metals have *low* values.

Learning Activity 5.3: The Structure of Water

Answer the following questions about the structure of water.

1. What part of the water molecule has a partial negative charge? A partial positive charge?

   Answer:

   The oxygen atom has a partial negative charge, while the hydrogen atoms have a partial positive charge.

2. Which element in water has the higher electronegativity value?

   Answer:

   Oxygen

3. Which part of the water molecule can form hydrogen bonds?

   Answer:

   Hydrogen bonds occur intermolecularly between hydrogen and oxygen atoms.

Learning Activity 5.4: Dissolving Compounds

1. What type of compound will dissociate in water? What type will not?

   Answer:

   Ionic compounds will dissociate; covalent compounds will not.

2. What part of the water molecule would be attracted to a solute ion such as Cl^-? Na^+?

   Answer:

   The hydrogen atoms. The oxygen atom.

3. How can you determine if the heat of reaction is endothermic or exothermic when looking at the chemical equation?

   Answer:

   If the heat appears on the reactant side, the reaction is endothermic. If the heat appears on the products side, the reaction is exothermic.

4. Write the equation for the dissolving of each of the following in water:

   a) $PbSO_{4(s)}$

      Answer:

      $PbSO_{4(s)} \rightarrow Pb^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$ ionic compound

   b) $C_6H_{12}O_{6(s)}$

      Answer:

      $C_6H_{12}O_{6(s)} \rightarrow C_6H_{12}O_{6(aq)}$ molecular compound

   c) $KBr_{(s)}$

      Answer:

      $KBr_{(s)} \rightarrow K^+{}_{(aq)} + Br^-{}_{(aq)}$ ionic compound

   d) $NaF_{(s)}$

      Answer:

      $NaF_{(s)} \rightarrow Na^+{}_{(aq)} + F^-{}_{(aq)}$ ionic compound

   e) $CH_3OH_{(l)}$

      Answer:

      $CH_3OH_{(l)} \rightarrow CH_3OH_{(aq)}$ molecular compound

   f) calcium chloride$_{(s)}$

      Answer:

      $CaCl_{2(s)} \rightarrow Ca^{2+}{}_{(aq)} + 2Cl^-{}_{(aq)}$ ionic compound

   g) $Na_2CO_{3(s)}$

      Answer:

      $Na_2CO_{3(s)} \rightarrow 2Na^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$ ionic compound

5. Describe what occurs during the process of an ionic solid dissolving into water. Include what happens to the solute and to the solvent.

   Answer:

   Solvent particles have to separate to make room for solute particles. Individual solute ions break away from the crystal. The negatively and positively charged ions become surrounded by solvent molecules. As this process continues, the ionic crystal dissolves.

Learning Activity 5.5: Predicting Solubility

Write a paragraph outlining your predictions for this lab. Which combinations of substances do you think will form a solution? Which combinations will not? On what do you base these predictions?

Answer:

Answers to this learning activity will vary, depending on your predictions. You are free to make any prediction, but make sure that your predictions are based on what you have learned in this lesson.

Learning Activity 5.6: Plotting a Solubility Curve

Step 1: Plot solubility on the *y*-axis (or vertical) using the units of grams solute/100 g H_20.

Step 2: Plot temperature in ^oC on the *x*-axis (or horizontal).

Step 3: Join the points together to create a curve. Note that you are NOT drawing a line of best fit.

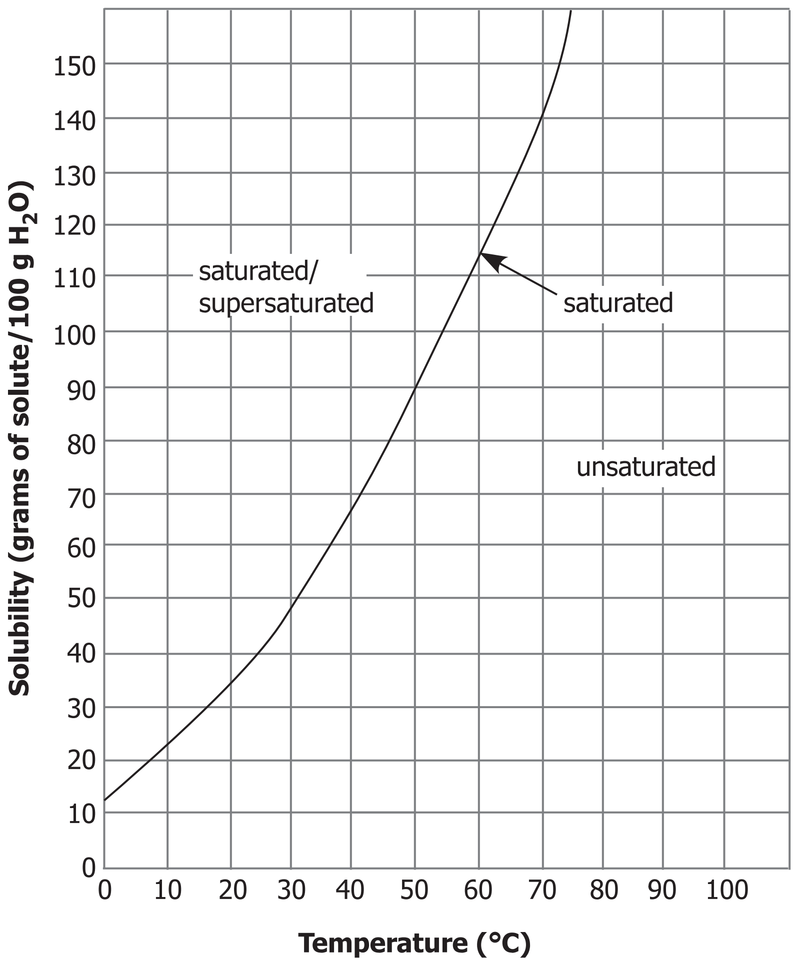
Step 4: Title the graph "Solubility Curve for KNO_3".

Step 5: Indicate on the graph where the following terms might apply: saturated, unsaturated, and supersaturated.

Temperature (°C)	Solubility (g solute/100 g H_2O)
0	12
20	34
40	68
60	112
70	140

Answer:

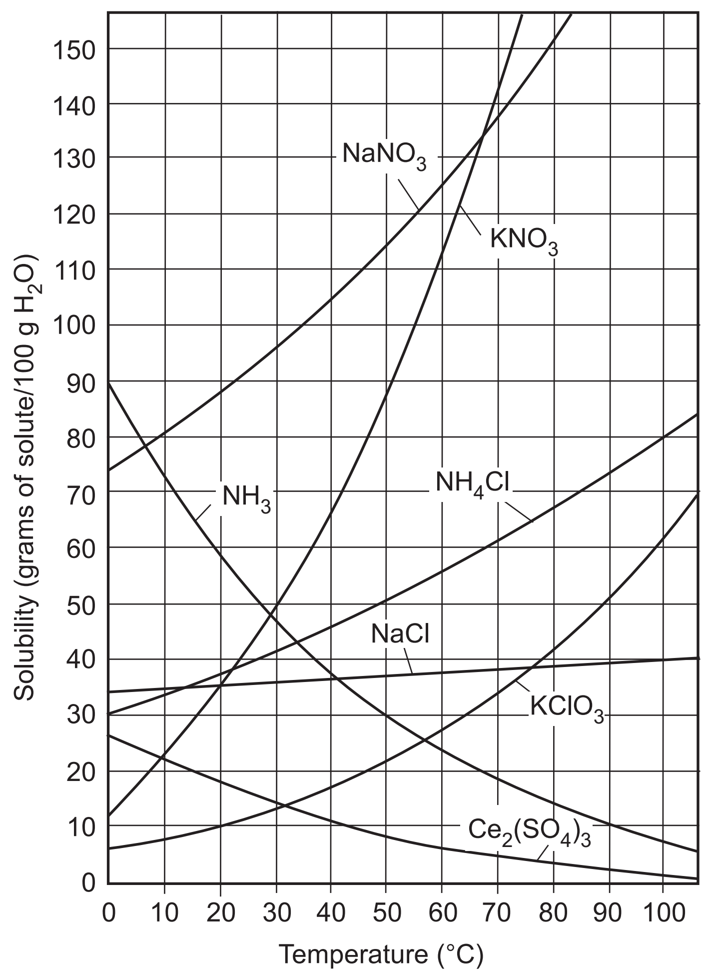
Solubility Curve for KNO_3



Learning Activity 5.7: Solubility and Temperature

Use the Solubility versus Temperature for Several Substances graph below to help solve the following problems (except for Questions 1 and 2, which do not require the graph). Show your calculations where necessary.

Solubility versus Temperature for Several Substances



1. Calculate the solubility of each of the following in g of solute/100 g of water.
   a) 0.250 kg dissolves in 1.2 L.

      Answer:

      $$\frac{250.\text{ g solute}}{1200\ \cancel{\text{g H}_2\text{O}}} \times 100\ \cancel{\text{g H}_2\text{O}} = 20.8 \approx 21\text{ g solute/100 g H}_2\text{O}$$

   b) 24.0 g dissolves in 280. g of water.

      Answer:

      $$\frac{24.0\text{ g solute}}{280.\ \cancel{\text{g H}_2\text{O}}} \times 100\ \cancel{\text{g H}_2\text{O}} = 8.57\text{ g solute/100 g H}_2\text{O}$$

2. Determine the solubility of the following in g of solute/L water.
   a) 261 g of a solid dissolves in 1510 mL of water.

      Answer:

      $$\frac{261\text{ g solute}}{1510\ \cancel{\text{mL}}\text{ H}_2\text{O}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} = 173\text{ g solute/L H}_2\text{O}$$

   b) 0.160 kg of a solid dissolves in 225 g of water.

      Answer:

      $$\frac{0.160\ \cancel{\text{kg}}\text{ solute}}{225\ \cancel{\text{g}}\text{ H}_2\text{O}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1000\text{ g}}{1\text{ L}} = 711\text{ g solute/L H}_2\text{O}$$

3. Determine the temperature of the following substances using the given solubilities and the graph provided.
   a) KNO_3 120 g/100 g

      Answer:

      63°C

   b) $NaNO_3$ 1200 g/L

      Answer:

      $$\frac{1200\text{ g NaNO}_3}{1\ \cancel{\text{L}}\ \cancel{\text{H}_2\text{O}}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{g}}} \times 100\ \cancel{\text{g}}\ \cancel{\text{H}_2\text{O}} = 120\text{ g NaNO}_3\text{/100 g H}_2\text{O}$$

      56° C

4. What is the solubility, in g/100 g water, of the following solutions at the specified temperature?
   a) $NaNO_3$ at 40°C

      Answer:

      105 g/100 g
   b) $Ce_2(SO_4)_3$ at 25°C

      Answer:

      16 g/100 g
   c) NH_3 at 30°C

      Answer:

      47 g/100 g
   d) NH_4Cl at 5°C

      Answer:

      32 g/100 g

5. How much more NH_4Cl can you dissolve in a litre of water at 60°C than at 20°C?

   Answer:

   At 60°C, the solubility is 55 g/100 g.

   At 20°C, the solubility is 37 g/100 g.

   55 g/100 g − 37 g/100 g = 18 g/100 g

$$\frac{18 \text{ g } NH_4Cl}{100 \text{ g } \cancel{H_2O}} \times \frac{1000 \text{ g } \cancel{H_2O}}{1 \text{ L}} = 180 \text{ g } NH_4Cl/1 \text{ L } H_2O$$

6. If you prepared a saturated solution of $NaNO_3$ at 80°C and then cooled it to 30°C, what would happen? Be specific.

   Answer:

   Solubility at 80°C is 151 g/100 g.

   Solubility at 30°C is 96 g/100 g

   151 g − 96 g = 55 g

   55 g of $NaNO_3$ will settle out for every 100 g of water.

7. At which temperature do $NaNO_3$ and KNO_3 have the same solubility? What about NaCl and NH_3?

   Answer:

   The same solubility occurs where the solubility curves cross on the graph. 67°C, 41°C.

8. How much water is needed to dissolve 65.0 g of NaNO3 at 35°C?

   Answer:

   At 35°C the solubility is about 100 g/100 g water.

$$65.0\ \cancel{\text{g NaNO}_3} \times \frac{100\ \text{g water}}{100\ \cancel{\text{g NaNO}_3}} = 65.0\ \text{g water}$$

9. What temperature is necessary to dissolve twice as much KNO_3 as can be dissolved at 30°C?

   Answer:

   At 30°C, the solubility of KNO_3 is 50 g/100 g water. The temperature at which its solubility is 100 g/100 g water is about 56°C.

10. If the solubility of a solid in water is 118 g/L, how much water would you need to dissolve a piece of the same solid with a mass of 45.0 g?

    Answer:

$$\frac{118\ \text{g solute}}{1\ \text{L H}_2\text{O}} = \frac{45.0\ \text{g solute}}{x\ \text{L H}_2\text{O}}$$

$$45.0\ \cancel{\text{g solute}} \times \frac{1\ \text{L H}_2\text{O}}{118\ \cancel{\text{g solute}}} = 0.381\ \text{L}\ \ \text{H}_2\text{O}$$

    0.381 L or 381 mL of water would be needed to dissolve 45 g of solute.

11. If 18.0 g of KNO_3 are dissolved in 15.0 mL of water at 100°C, at what temperature will the solid begin to settle out?

    Answer:

$$\frac{18\ \text{g KNO}_3}{15\ \cancel{\text{g H}_2\text{O}}} \times 100\ \cancel{\text{g H}_2\text{O}} = 120\ \text{g KNO}_3/100\ \text{g H}_2\text{O}$$

    At about 62°C the solubility of KNO_3 is 120 g/100 g water. KNO_3 will begin settling out at 62°C.

12. If 40.0 g of KNO_3 is added to 50.0 mL of water at 40°C, will it all dissolve? If not, how much would be leftover? If you raised the temperature to 45°C, will it all dissolve? Give evidence.

Answer:

$$\frac{40.0\ \text{g KNO}_3}{50.0\ \cancel{\text{g H}_2\text{O}}} \times 100\ \cancel{\text{g H}_2\text{O}} = 80.0\ \text{g KNO}_3/100\ \text{g H}_2\text{O}$$

At 40°C, the solubility of KNO_3 is about 66 g/100 g water, so not all of the solute will dissolve.
80.0 g − 66.0 g = 14.0 g
14.0 g will not dissolve at 40°C.
At 45°C, the solubility is about 79 g/100 g water. Not all of the solute will dissolve.

13. If 142 g of NH_4Cl are dissolved in 350. mL of water at 55°C, is the solution saturated?

Answer:

$$\frac{142\ \text{g NH}_4\text{Cl}}{350.\ \cancel{\text{g H}_2\text{O}}} \times 100\ \cancel{\text{g H}_2\text{O}} = 40.6\ \text{g NH}_4\text{Cl}/100\ \text{g H}_2\text{O}$$

The solubility of NH_4Cl at 55°C is about 53 g/100 g water. The solution is therefore unsaturated.

Learning Activity 5.8: The Effect of Pressure on Solubility

Complete the following Learning Activity to help you understand the effect of pressure on the solubility of a gas in a liquid. Use a bottle or can of your favourite soft drink to help you answer these questions.

1. Is this soft drink a solution? How do you know?

   Answer:

   All examples of soda pop and soft drinks are solutions. They are transparent, they do not settle out upon standing, and only one phase (liquid) is visible.

2. Look on the label. What are the solutes? What is the solvent? How do you know?

   Answer:

   The solutes may vary slightly, but all sodas contain carbon dioxide (carbonated water), sugar, citric acid, etc. The solvent is water (carbonated water, more specifically), as the ingredients present in the greatest amount are always listed first.

3. Are the solutes and solvent polar or non-polar?

   Answer:

   Since water is polar, most of the solutes will also be polar as they are soluble and miscible in water. However, carbon dioxide is non-polar and is held into water under pressure.

4. Why does the drink make a sound when you open it?

   Answer:

   When the drink is opened, the pressure of the undissolved gas trapped inside the can drops to the pressure of the atmosphere around the can. As a result, the undissolved gas trapped above the liquid escapes. The sound you hear is made by the escaping gas. Read on to find out how the solubility of the gas in the liquid is affected by this pressure change!

Learning Activity 5.9: Colligative Properties

1. What factor determines how the vapour pressure, freezing point, and boiling point of a solution differ from those of a pure substance?

   Answer:

   The number of solute particles dissolved in the solvent affects all the above.

2. Would a dilute (weak) or a concentrated (strong) solution of saltwater have a higher boiling point? Why?

   Answer:

   The concentrated solution would have the higher boiling point, since boiling point elevation is proportional to the number of solute particles dissolved in the solvent.

3. Equal quantities of KI and MgI_2 are dissolved into 1 L of water. Which solution will have the higher boiling point, and why?

   Answer:

   The MgI_2 will have the higher boiling point. When it dissolves, it forms three ions (Mg^{2+}, I^-, and I^-), unlike the two ions formed by potassium iodide (K^+ and I^-). Since there are more particles of solute, the boiling point will increase more for the magnesium iodide.

4. What are the three colligative properties of solutions?

   Answer:

   The three colligative properties of solutions are: vapour pressure lowering, freezing point depression, and boiling point elevation.

Learning Activity 5.10: Solving Concentration Problems

Solve the following problems, showing your calculations.

1. Calculate the concentration of a 200.0 mL solution that contains 0.250 moles of solute.

   Answer:

   $$\text{Concentration} = \frac{\text{moles}}{\text{Litres}} = \frac{0.250\ \text{mol}}{0.2000\ \text{L}} = 1.25\ \text{mol/L}$$

2. Find the concentration of a solution that contains 1.45 moles dissolved in 2.30 L of solution.

   Answer:

   $$\text{Concentration} = \frac{\text{moles}}{\text{Litres}} = \frac{1.45\ \text{mol}}{2.30\ \text{L}} = 0.630\ \text{mol/L}$$

3. How many moles of NaOH would be needed to make 50.0 mL of a 0.750 mol/L solution?

   Answer:

   $$\text{moles} = \left(\frac{0.750\ \text{mol}}{\cancel{\text{L}}}\right)\left(0.0500\ \cancel{\text{L}}\right) = 0.0375\ \text{mol}$$

4. What mass of $AgNO_3$ would be needed to make 250.0 mL of a 1.50 mol/L solution?

   Answer:

   $$AgNO_3 = 169.9\ \text{g/mol}$$

   $$\text{mass} = 0.2500\ \cancel{\text{L}} \times \frac{1.50\ \cancel{\text{mol}}}{1\ \cancel{\text{L}}} \times \frac{169.9\ \text{g}}{1\ \cancel{\text{mol}}} = 63.7\ \text{g}$$

5. What mass of $CaCO_3$ would be needed to make 20.0 mL of a 0.400 mol/L solution?

   Answer:

   $$CaCO_3 = 100.1\ \text{g/mol}$$

   $$\text{mass} = 0.0200\ \cancel{\text{L}} \times \frac{0.400\ \cancel{\text{mol}}}{1\ \cancel{\text{L}}} \times \frac{100.1\ \text{g}}{1\ \cancel{\text{mol}}} = 0.801\ \text{g}$$

6. How much solution would be needed to dissolve 50.0 g of K_2SO_4 to make a 0.500 mol/L solution?

   Answer:

   $K_2SO_4 = 174.3 \text{ g/mol}$

   $$\text{volume} = 50.0\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{174.3\ \cancel{\text{g}}} \times \frac{1\ \text{L}}{0.500\ \cancel{\text{mol}}} = 0.574\ \text{L}$$

7. What volume of solution would be required to dissolve 18.04 g of aluminum sulfide to make a 0.160 mol/L solution?

   Answer:

   $Al_2S_3 = 150.3 \text{ g/mol}$

   $$\text{volume} = 18.04\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{150.3\ \cancel{\text{g}}} \times \frac{1\ \text{L}}{0.160\ \cancel{\text{mol}}} = 0.750\ \text{L}$$

8. What mass of sodium sulfate would be needed to make 50.0 mL of a 0.150 mol/L solution?

   Answer:

   $Na_2SO_4 = 142.1 \text{ g/mol}$

   $$\text{mass} = 0.0500\ \cancel{\text{L}} \times \frac{0.150\ \cancel{\text{mol}}}{1\ \cancel{\text{L}}} \times \frac{142.1\ \text{g}}{1\ \cancel{\text{mol}}} = 1.06\ \text{g}$$

Learning Activity 5.11: Preparing a Solution

1. When preparing a solution, why can you not obtain the required volume first and *then* add the mass of solute?

   Answer:

   If you measure the total volume of solvent and then mass the solute, you will end up with a greater volume than you want to have. This, in turn, will affect the desired concentration of the solution you are trying to prepare.

2. Why is it necessary to shake or swirl the solution well upon adding the final volume of solvent?

   Answer:

   Shaking the contents of your flask will ensure that all of the solute dissolves completely, and ensures the homogeneity of the solution you have prepared.

3. You prepare 250.0 mL of a solution using 11.5 g of potassium bromide, KBr. What is the concentration of the prepared solution?

   Answer:

$$\text{Moles of KBr} = 11.5\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{119.0\ \cancel{\text{g}}} = 0.0966\ \text{mol KBr}$$

$$\text{Concentration} = \frac{0.0966\ \text{mol KBr}}{0.2500\ \text{L}} = 0.386\ \text{mol/L}$$

Learning Activity 5.12: Dilution Calculations, Part 1

Solve the following dilution problems, showing your calculation for each one.

1. Find the final concentration when 600.0 mL of 6.00 mol/L solution has 200.0 mL of water added to it.

   Answer:

   $V_2 = V_1$ + volume added = 600.0 mL + 200.0 mL = 800.0 mL = 0.8000 L

   $$C_2 = \frac{C_1V_1}{V_2} = \frac{(6.00 \text{ mol/L})(0.6000 \cancel{\text{ L}})}{0.8000 \cancel{\text{ L}}} = 4.50 \text{ mol/L}$$

2. A 0.500 mol/L solution has a volume of 8.00 L. The initial concentration was 2.50 mol/L. What was the initial volume?

   Answer:

   $$V_1 = \frac{C_2V_2}{C_1} = \frac{(0.500 \cancel{\text{ mol/L}})(8.00 \text{ L})}{2.50 \cancel{\text{ mol/L}}} = 1.60 \text{ L}$$

3. You start with a solution that is 0.800 mol/L and 70.0 mL of water. You need to prepare a 0.300 mol/L solution. What is the final volume of the solution?

   Answer:

   $$V_2 = \frac{C_1V_1}{C_2} = \frac{(0.800 \cancel{\text{ mol/L}})(0.0700 \text{ L})}{0.300 \cancel{\text{ mol/L}}} = 0.187 \text{ L}$$

4. You have 500.0 mL of 3.00 mol/L sodium chloride solution. The solution is diluted to 580.0 mL. What is the final concentration?

   Answer:

   $$C_2 = \frac{C_1V_1}{V_2} = \frac{(3.00 \text{ mol/L})(0.5000 \cancel{\text{ L}})}{0.5800 \cancel{\text{ L}}} = 2.59 \text{ mol/L}$$

5. 8.00 L of a 0.300 mol/L acid must be diluted to 0.0100 mol/L before it can safely be put into the sewage system. What is the final volume of the solution, and how much water must be added?

   Answer:

$$V_2 = \frac{C_1V_1}{C_2} = \frac{(0.300 \cancel{\text{mol/L}})(8.00\ \text{L})}{0.0100 \cancel{\text{mol/L}}} = 240.\ \text{L}$$

   Water added = $V_2 - V_1$ = 240. L − 8.00 L = 232 L of water added.

6. 60.0 mL of 6.00 mol/L sulfuric acid is diluted to 5.00 L. What is the final concentration?

   Answer:

$$C_2 = \frac{C_1V_1}{V_2} = \frac{(6.00\ \text{mol/L})(0.0600\ \cancel{\text{L}})}{5.00\ \cancel{\text{L}}} = 0.0720\ \text{mol/L}$$

7. Describe how to make 250.0 mL of a 0.100 mol/L solution of hydrochloric acid (HCl) from a solution that is 11.8 mol/L.

   Answer:

$$V_1 = \frac{C_2V_2}{C_1} = \frac{(0.100\ \cancel{\text{mol/L}})(0.2500\ \text{L})}{11.8\ \cancel{\text{mol/L}}} = 0.00212\ \text{L}$$

   0.00212 L or 2.12 mL of HCl to a 250 mL volumetric flask. Fill the flask to the mark with water and swirl the solution.

8. If 45.0 mL of stock HCl was required to make 150.0 mL of a 3.48 mol/L solution, calculate the concentration of the original stock solution.

   Answer:

$$C_1 = \frac{C_2V_2}{V_1} = \frac{(3.48\ \text{mol/L})(0.1500\ \cancel{\text{L}})}{0.0450\ \cancel{\text{L}}} = 11.6\ \text{mol/L}$$

Learning Activity 5.13: Preparing a Solution

1. How is preparing a *dilution* different from preparing a solution?

   Answer:

   Preparing a dilution does not involve massing out the solute, like preparing a solution does.

   Preparing a dilution involves calculating the required amount of solution, while preparing a solution does not require this calculation.

2. How is preparing a dilution *similar* to preparing a solution?

   Answer:

   Preparing both involves precise measurements of volume.

   Preparing solutions and dilutions require adding the solvent to the mark on the volumetric flask as the final step, preferably with a pipette or eyedropper.

Learning Activity 5.14: Dilution Calculations, Part 2

1. What is the final concentration when a 1.0 L of a 0.10 mol/L solution is mixed with 1.0 L of a 1.0 mol/L of the same solution?

   Answer:

$$\text{Moles of solution 1: } \frac{0.10 \text{ mol}}{\cancel{\text{L}}} \times 1.0 \ \cancel{\text{L}} = 0.10 \text{ mol}$$

$$\text{Moles of solution 2: } \frac{1.0 \text{ mol}}{\cancel{\text{L}}} \times 1.0 \ \cancel{\text{L}} = 1.0 \text{ mol}$$

$$\text{Final concentration} = \frac{\text{moles 1} + \text{moles 2}}{\text{volume 1} + \text{volume 2}} = \frac{0.10 \text{ mol} + 1.0 \text{ mol}}{1.0 \text{ L} + 1.0 \text{ L}}$$

$$= 0.55 \text{ mol/L}$$

2. What is the final concentration when 400.0 mL of a 0.050 mol/L HCl solution is mixed with exactly 600.0 mL of a 0.020 mol/L HCl solution?

Answer:

$$\text{Moles of solution 1}: \frac{0.050 \text{ mol}}{\cancel{\text{L}}} \times 0.4000 \ \cancel{\text{L}} = 0.020 \text{ mol}$$

$$\text{Moles of solution 2}: \frac{0.020 \text{ mol}}{\cancel{\text{L}}} \times 0.6000 \ \cancel{\text{L}} = 0.012 \text{ mol}$$

$$\text{Final concentration} = \frac{\text{moles 1} + \text{moles 2}}{\text{volume 1} + \text{volume 2}} = \frac{0.020 \text{ mol} + 0.012 \text{ mol}}{0.400 \text{ L} + 0.600 \text{ L}}$$

$$= 0.032 \text{ mol/L}$$

Learning Activity 5.15: Water Treatment Methods

1. Research then summarize the following water treatment methods:

a) Fluorination

Answer:

Fluorination is the addition of fluoride to public water supplies to reduce tooth decay.

b) Distillation

Answer:

Distillation is a process that removes almost all impurities from water (or other liquids). The water is boiled and the steam is condensed into a clean container, leaving most if not all of the solid contaminants behind.

c) Aeration

Answer:

Aeration is the process of adding oxygen to waste water in order to encourage the growth of certain bacteria that treat the water. This is often done in a lagoon where the waste water is stored.

d) Reverse Osmosis

Answer:

Reverse osmosis (RO) is the process of using pressure to force water through a series of filters and membranes. The impurities are left on one side of the membranes, while the pure water passes through. The filters and membranes might include:

- one or more sediment filters to trap larger particles
- one or more activated carbon filters, which trap organic chemicals
- an RO filter
- an ultra-violet lamp for disinfecting microbes

2. What is your local water source? See the Manitoba Water Stewardship Branch website at <www.gov.mb.ca/waterstewardship.odw/index.html> for information on your community water supply.

   Answer:

   Possible answers include wells, reservoirs, etc.

3. What does the term potable water mean? Give an example of non-potable water.

   Answer:

   It means that the water is fit for human consumption. Examples of non-potable water include: water in the bathrooms on trains and airplanes, lake and river water that is untreated, sea water, rain water, and others.

Notes

Grade 11 Chemistry (30S)

Module 6: Organic Chemistry

Module 6: Organic Chemistry

Introduction

How are fossils and chemistry related? In this module, you will answer that question and many others related to organic chemistry. What does the term *organic* mean? What are some sources of hydrocarbons?

Next, you will learn how to name, draw, and construct the structural formulae for many groups of hydrocarbons. Your friends will be impressed when you can name the acid that makes an ant bite sting!

Molecular Model Kit

In order to construct the structural formulae of hydrocarbons, you will need access to a molecular model kit. You will start using this kit in Lesson 3. If you have not yet purchased it, you can do so now by contacting the Learning Resource Centre at 1-866-771-6822 or go online at www.manitobalrc.mb.ca. Ask for item number 7765.

Assignments in Module 6

When you have completed the assignments for Module 6, submit your completed assignments to the Distance Learning Unit either by mail or electronically through the learning management system (LMS). The staff will forward your work to your tutor/marker.

Lesson	Assignment Number	Assignment Title
1	Assignment 6.1	Origins and Major Sources of Hydrocarbons
2	Assignment 6.2	Aliphatic Hydrocarbons
3	Assignment 6.3	Naming and Drawing Alkanes
4	Assignment 6.4	Alkane Isomers
5	Assignment 6.5	Naming and Drawing Alkenes
6	Assignment 6.6	Naming and Drawing Alkynes
7	Assignment 6.7	Researching an Aromatic Compound
8	Assignment 6.8	Naming and Drawing Alcohols
9	Assignment 6.9	Drawing Carboxylic Acids
10	Assignment 6.10	Naming and Drawing Esters
11	Assignment 6.11	Chemistry and Our Quality of Life
12	Assignment 6.12	Investigating an Issue in Organic Chemistry

As you work through this course, remember that your learning partner and your tutor/marker are available to help you if you have questions or need assistance with any aspect of the course.

Writing Your Final Examination

You will write the final examination when you have completed Module 6 of this course. The final examination is based on Modules 1 to 6, and is worth 30 percent of your final mark in the course. Please note that 80–85 percent of the final examination is concentrated on Modules 4 to 6. To do well on the final examination, you should review all the work you complete in Modules 1 to 6, including all the learning activities and assignments. You will write the final examination under supervision.

NOTES

Lesson 1: The Chemistry of Carbon (1.5 hours)

Lesson Focus

SLO C11-5-01: Compare and contrast inorganic and organic chemistry.
Include: the contributions of Friedrich Wöhler and the overturn of vitalism

SLO C11-5-02: Identify the origins and major sources of hydrocarbons and other organic compounds.
Include: natural and synthetic sources

Lesson Introduction

Recently, the term "organic" has been used to describe food grown without the aid of synthetic fertilizers, pesticides, and other chemical enhancements. In chemistry, organic compounds are compounds containing the element carbon. Plastics, gasoline, drugs, natural gas, car tires, and the clothes you wear are just some examples of the many things that are made of organic compounds. Even the foods you eat are largely made of organic compounds—carbohydrates, fats, and proteins. Thus, organic compounds are found everywhere. In this module, you will compare organic and inorganic chemistry. You will also study several types of organic compounds and their sources.

Organic or Inorganic?

Chemists divide all known compounds into two classes, organic and inorganic. Organic compounds are all carbon-containing compounds, *with the exception of* CO, CO_2, CS_2, carbonates (CO_3^{2-}), bicarbonates (HCO_3^-), cyanides (CN^-), and carbides (i.e., CaC_2). Inorganic compounds include those exceptions and all other compounds. Of the over 13,000,000 known compounds, only about 100,000 are classified as inorganic. **Organic chemistry**, the focus of this lesson, is the study of the structure, composition, properties, and reactions of organic compounds.

Friedrich Wöhler

In the early 1700s, the term **organic** implied that the source of the substance was a living organism (that is, plants and animal sin nature). In fact, it was believed that only living organisms could make the carbon compounds found in their cells. Organic chemistry at that time referred only to the study of these carbon-based compounds found in living cells. It was believed that living organisms had a **"vital force"** that allowed them to produce such compounds. Because of this "vital force," organic compounds could only be isolated or synthesized from living, or previously living, things—until Friedrich Wöhler came along.

Friedrich Wöhler (1800–1882) was a German chemist who, at the age of 20, began attending Marburg University in Germany and worked in the laboratory of the chemist Leopold Gmelin. Wöhler initially intended to become a physician, but Gmelin convinced him to devote himself to the study of chemistry. Wöhler spent much of his early days in chemistry isolating various elements such as aluminum (Al), beryllium (Be), boron (B), silicon (Si), and titanium (Ti). In 1828, he attempted to prepare ammonium cyanate (NH_4OCN) by means of a double decomposition reaction in a solution of ammonium chloride (NH_4Cl) and silver cyanate (AgOCN). Both of these compounds were considered to be 'inorganic.'

Wöhler was not able to produce ammonium cyanate as he had anticipated. When he analyzed the products of his reaction, he found that he had produced urea (CH_4N_2O), an organic compound found in urine!

$$NH_4Cl + AgOCN \rightarrow AgCl + \underset{\text{Urea}}{CH_4N_2O}$$

At that point in history, urea had only been extracted from animal kidneys. This was an amazing discovery, since it was the first time that inorganic compounds without the "vital force" had been used to synthesize an organic compound with "vital force."

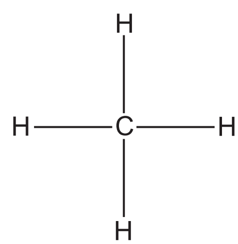
Wöhler's chance discovery was the beginning of the end of the "vital force" theory. Within a few years of this event, when acetic acid and several other organic compounds had been prepared from inorganic materials, the validity of the "vital force" theory was questioned. As time passed, even more organic compounds were synthesized from inorganic materials. It became obvious that it was not necessary for all organic compounds to be associated with living organisms, but rather that all organic compounds contained the element carbon. Chemists now simply say that organic compounds are compounds containing carbon. Today, organic chemistry includes the study of *all* carbon-based compounds, regardless of their origin.

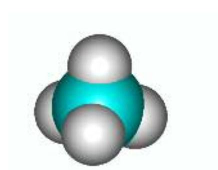
What is a Hydrocarbon?

It is likely that you use at least one hydrocarbon every day. A hydrocarbon is the simplest form of an organic compound, as it contains only carbon and hydrogen. It is noteworthy that most hydrocarbons are extracted from petroleum. You may think that such a simple compound could not be very useful, but this is the wrong assumption to make! There are millions of different hydrocarbons, some which are routinely used in everyday life such as:

- The gasoline used in vehicles, lawn mowers, boats, etc.
- The diesel fuel used in some trucks, buses, and cars.
- The solid coal that is burned for heat.
- The natural gas used for home heating and barbeques.

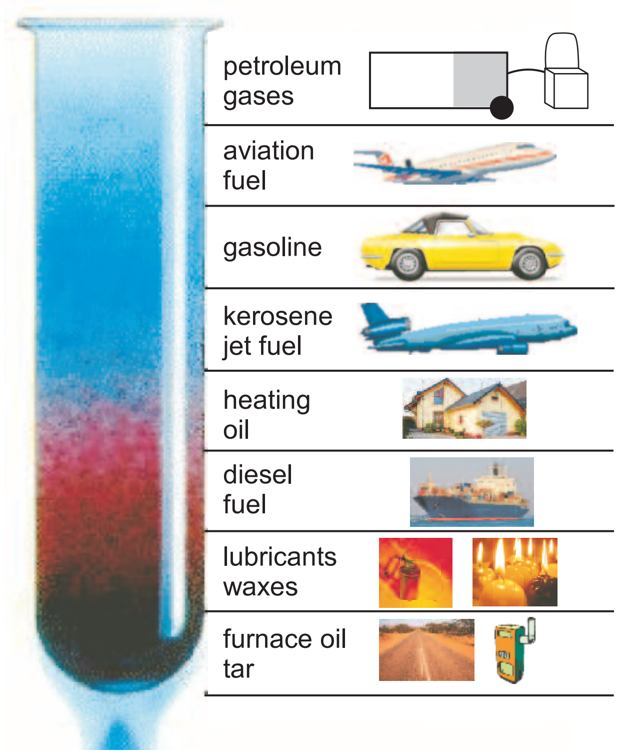
All hydrocarbons are non-polar molecules. As a result, they are insoluble in water. The non-polar nature of the hydrocarbons results in very low intermolecular forces, and very low melting and boiling points in relation to their mass. Hydrocarbons will burn in oxygen to form carbon dioxide and water. Methane, shown below, is the simplest hydrocarbon:





Sources of Organic Compounds

In Grade 10 Science, while studying the carbon cycle, you may have learned that naturally occurring organic compounds result from the decay of prehistoric animals and vegetation. Natural hydrocarbons are formed from the combination of immense *pressure* on collected sediments along with *heat* from the Earth's core on organic material. The decaying process is accomplished by *bacteria*, contributing to the formation of organic material. Hydrocarbon fuels from this process are generally called fossil fuels or petroleum products. The word **petroleum** is derived from the Latin roots "*petra*" (meaning rock) and "*oleum*" (for oil). **Crude oil** is the name given to the unprocessed petroleum that is extracted from the ground. The refining of crude oil involves several steps, two of which are briefly described below.

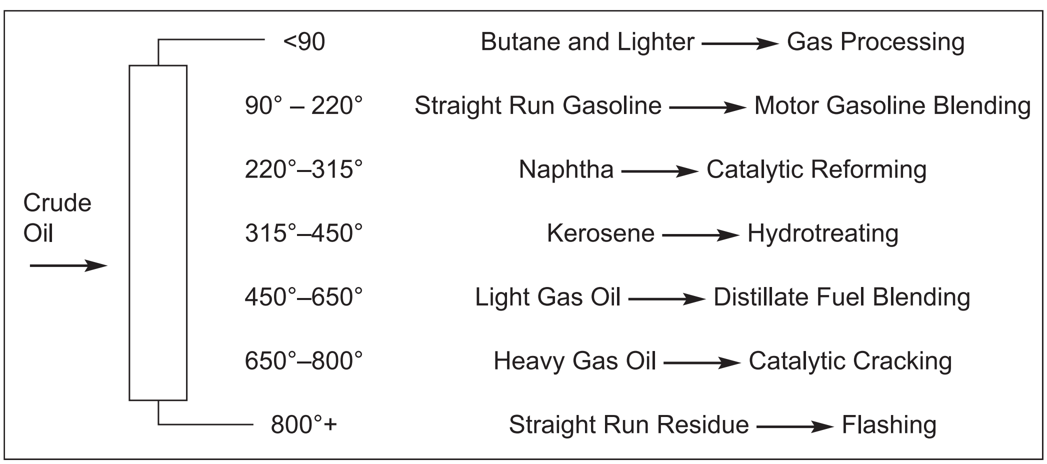


Distillation

When crude oil is extracted from a well, it contains a mixture of hydrocarbon compounds and relatively small quantities of other materials such as oxygen, nitrogen, sulfur, salt, and water. In the refinery, most of these non-hydrocarbon substances are removed and the oil is broken down into its various components, so that it can be blended into useful products.

These compounds can be separated based on their boiling points using a process called **fractional distillation** (or **fractionation**). Remember that, generally speaking, *the larger the compound, the higher its boiling point*. When the crude oil is heated to 350–400ºC, both the vapour and liquid are piped into the distilling column. While the liquid falls to the bottom, the vapour rises, passing through a series of perforated trays (called sieve trays). *Heavier hydrocarbons condense more quickly and settle on lower trays and lighter hydrocarbons remain as a vapour longer and condense on higher trays*.

Liquid fractions are drawn from the trays and removed for further purification and refining. Over time, much of the oil will boil and vaporize. The oil that does not vaporize is called residue, and it consists of compounds containing 20 or more carbon atoms. Residue drawn from the bottom of the distilling column may be burned as fuel or processed into lubricating oils, waxes, and bitumen. To recover additional heavy distillates from this residue, it may be piped to a second distillation column where the process is repeated under vacuum. This **vacuum distillation** allows heavy hydrocarbons with boiling points of 450°C and higher to be separated.



The diagram above shows that light gases (such as methane, ethane, propane, and butane) pass out the top of the distillation column. Straight run gasoline (gasoline comprised of only natural ingredients from crude oil or natural gas) is formed in the top trays, kerosene and gas oils in the middle, and fuel oils at the bottom.

Cracking

Cracking is a process where larger molecules are broken down into more useable small molecules. More specifically, cracking processes break down heavier hydrocarbons (oils with higher boiling points) into lighter products such as gasoline and diesel. These processes include catalytic cracking, thermal cracking, and hydrocracking.

In **catalytic cracking**, a catalyst (such as a sample of small pellets of silica) is added to petroleum in the absence of oxygen. The products of cracking are used in the making of gasoline and starting molecules for synthetic petroleum products. **Thermal cracking** uses heat to break down the residue from vacuum distillation. The lighter elements produced from this process can be made into distillate fuels and petrol. **Hydrocracking** is catalytic cracking in the presence of hydrogen, whereby the extra hydrogen saturates, or hydrogenates, the chemical bonds of the cracked hydrocarbons. Hydrocracking is also a treating process, because the hydrogen combines with contaminants such as sulfur and nitrogen, allowing them to be removed.

Finally, **synthetically produced hydrocarbons** are constructed by starting with a petrochemical compound and adding to it to create longer hydrocarbon chains. Many synthetic products are more stable at higher temperatures, and are very insoluble, making them excellent lubricants. There are many patented methods of producing petroleum-like products. For example, **tar sands** (also referred to as **oil sands**) are a combination of clay, sand, water, and bitumen. Alberta's tar sands are mined and processed to extract the oil-rich **bitumen**, which is then refined into oil.

The bitumen in tar sands is heavy and viscous; therefore, it cannot be pumped from the ground in its natural state. Instead, tar sand deposits are mined using strip mining or open pit techniques, or the oil is extracted by underground heating with additional refining. Approximately 75% of the bitumen present can be recovered from tar sand. After oil extraction, the spent sand and other materials are then returned to the strip mine, with the hope that the land will eventually be reclaimed and the ecosystem returned to its former state. About two tons of tar sands and enormous amounts of water are required to produce one barrel of oil, placing enormous stress on an ecosystem. You will have an opportunity to learn more about Canada's tar sands and explore the views of different stakeholders with respect to this important resource in Lesson 12.

Learning Activity 6.1

Organic Chemistry Introduction

1. Explain the two ways that chemists classify substances.
2. How did Friedrich Wöhler change the way scientists perceived the nature of organic compounds?
3. Define the term “organic compound,” according to the findings of Friedrich Wöhler.
4. Name and briefly describe the two steps involved in the refining of crude oil.
5. List three major sources of hydrocarbons.
6. List three distinguishing properties of hydrocarbons.
7. What is petroleum?

Lesson Summary

In this lesson, you learned about the difference between inorganic and organic chemistry. You were introduced to the contributions of Friedrich Wöhler and how his work furthered our understanding of organic compounds. Finally, you learned about the major sources of hydrocarbons and other organic compounds. In the next lesson, you will continue your study of hydrocarbons by comparing and contrasting alkanes, alkenes, and alkynes.

NOTES

Assignment 6.1

Origins and Major Sources of Hydrocarbons *(6 marks)*

1. Identify each of the following as organic or inorganic. *(6 marks)*

   a) $NaHCO_3$

   ____

   b) CH_4

   ____

   c) $HC_2H_3O_2$

   ____

   d) HCN

   ____

   e) Ca_2C

   ____

   f) C_2H_2

   ____

Notes

Lesson 2: Hydrocarbons (2 hours)

Lesson Focus

SLO C11-5-03: Describe the structural characteristics of carbon.
Include: bonding characteristics of the carbon atom in hydrocarbons (single, double, triple bonds)

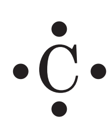
SLO C11-5-04: Compare and contrast the molecular structures of alkanes, alkenes, and alkynes.
Include: trends in melting points and boiling points of alkanes only

Lesson Introduction

In the previous lesson, you were introduced to hydrocarbons. In this lesson, you will continue your study of hydrocarbons, first by describing the structural characteristics of carbon. Next, you will learn about three classifications of hydrocarbons—alkanes, alkenes, and alkynes.

The Chemistry of Carbon

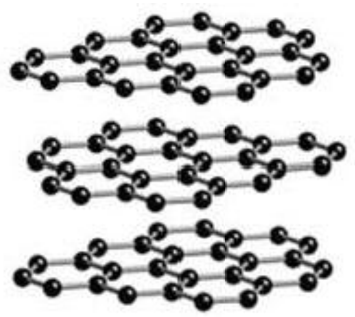
There are literally millions of carbon compounds—why are there so many? This astounding number of compounds is a result of the unique characteristics of the carbon atom. Carbon has four valence electrons; that is, four electrons in its outer shell. Therefore, in its effort to fill its valence shell with eight electrons, carbon will always form four covalent bonds, or share four pairs of electrons. You may recall drawing the Lewis Dot Diagram of carbon in Grade 10 Science, which illustrates the combining capacity of this element:



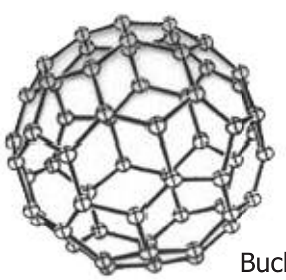
Because carbon is a small atom, it can share 1, 2, or 3 pairs of electrons with other carbon atoms to form long chains or ring structures. Carbon atoms are unique in that they will even form bonds with other carbons while they are covalently bound to other atoms like hydrogen, oxygen, nitrogen, and the halogens.

Allotropes of Carbon

Carbon is also unique in that it is an element that can exist in two or more different forms, or **allotropes**, of that element. In each different allotrope, the element's atoms are bonded together in a different manner. For example, in **diamond**, the carbon atoms are bonded together in a tetrahedral lattice arrangement, while in **graphite**, the carbon atoms are bonded together in sheets of a hexagonal lattice, as shown below:



If a sheet of graphite were to form into a ball so that the edges would meet, the bonding capacity of all atoms would be satisfied. This arrangement of atoms looks much like the geodesic domes designed by the architect-engineer, R. Buckminster Fuller. You may recognize this structure of interlocking hexagons and pentagons, as it is identical to that of a soccer ball. Because this model was inspired by the geodesic dome, scientists named this allotrope of carbon **buckminsterfullerene**, which is also known as the **buckyball** or **fullerene**.

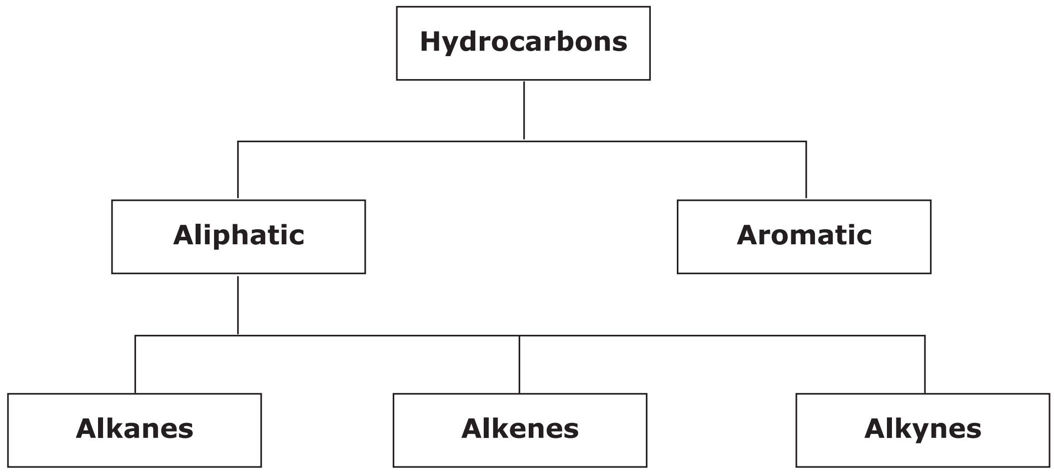


Buckminsterfullerene

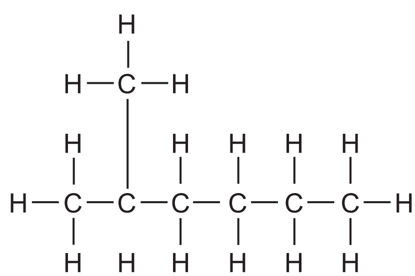
The covalent bonds formed by carbon atoms are extremely strong. This is evident by the hardness of diamonds, which are the hardest substances in nature because of the strength of their carbon-carbon bonds.

Classifying Hydrocarbons

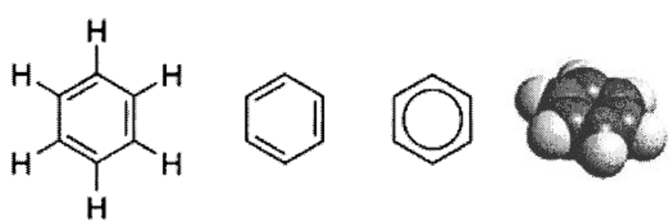
Hydrocarbons can be divided into several groups, as shown in the diagram below:



Aliphatic hydrocarbons have carbon atoms arranged in straight or branched chains. They can also form ringed structures, although this lesson will focus on the straight and branched chains such as 2-methylhexane (shown below).



Aromatic hydrocarbons, like the ones illustrated below, contain a benzene ring. You will learn more about the aromatic hydrocarbons later in this module.



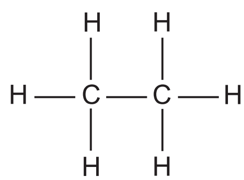
Aliphatic hydrocarbons do not contain benzene ring structures and can be divided into alkanes, alkenes, and alkynes.

Alkanes

Alkanes contain only single carbon-carbon bonds, where the carbon atoms share only one pair of electrons. The general molecular formula for alkanes is:

$$C_nH_{2n+2}$$

In this formula, *n* represents the number of carbon atoms (that is, 1, 2, 3, ….), thus producing formulas like C_2H_6 and C_3H_8. Ethane, shown below, is an example of an alkane:

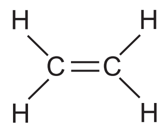


Those compounds whose carbon-carbon bonds are all single bonds are referred to as being **saturated**. This means that each carbon is bound to four atoms, the maximum number possible, as shown in the example above. Alkanes in which the carbon atoms form long chains are called **normal, straight-chain**, or **unbranched hydrocarbons**. *As the molecular weights of normal alkanes increase, so do their boiling and melting points.*

Boiling and Melting Points of Selected Normal Alkanes				
Name	**Formula**	**Molar Masses**	**Melting Point (°C)**	**Boiling Point (°C)**
Methane	CH_4	16	–183	–162
Propane	C_3H_8	44	–187	–42
Hexane	C_6H_{14}	86	–95	68
Octane	C_8H_{18}	114	–57	126
Decane	$C_{10}H_{22}$	142	–30	174

Alkenes

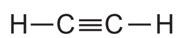
Alkenes are a group of aliphatic hydrocarbons that contain one or more double carbon-carbon bonds. The carbons share two pairs of electrons, resulting in a double bond, as illustrated with ethylene below. Note that there are still four bonds around each carbon, as you would expect.



The general molecular formula for alkenes is C_nH_{2n}.

Alkynes

Alkynes are hydrocarbons that contain one or more triple carbon-carbon bond. The general molecular formula for alkenes is C_nH_{2n-2}. Acetylene, shown below, is an example of an alkyne. Again, you will observe four bonds around each carbon.

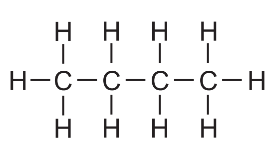


Organic compounds that contain double and triple carbon-carbon bonds are called **unsaturated hydrocarbons** because they possess fewer than the maximum number of hydrogens in their structure.

Molecular and Structural Formulas

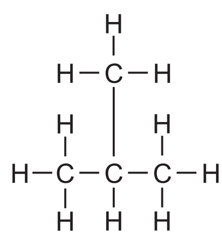
The **molecular formula** of an organic compound shows the kind of atoms and the number of each kind of atom in the compound. For example, the molecular compound for butane, used in disposable lighters, is C_4H_{10}. Butane has four carbon atoms and ten hydrogen atoms.

The **structural formula** of an organic compound shows the arrangement of the atoms found in the molecular formula. One possible structural formula for C_4H_{10} is butane:



Looking at this structure you can see that there are in fact four carbon atoms and ten hydrogen atoms in butane. Each carbon has four bonds and all hydrogens have one. Remember that carbon must always form four bonds. Therefore, when drawing structural formulas, always ensure that every carbon has four bonds. Hydrogen can only make one bond, since it only needs one electron to complete its valence shell. As such, hydrogen can only share one pair of electrons.

You may have noticed that the above representation of butane was noted as a *possibility*. This implies that there is another way to arrange four carbons and ten hydrogens. The arrangement of 2-methylpropane below still ensures that each of the four carbons creates four bonds and each of the ten hydrogens creates one bond.



Note that the structural formula shows the arrangement of atoms in a compound, while the molecular formula does not. Hydrocarbons with the same molecular formula but different arrangements of the atoms are called **structural isomers**.

Condensed Structural Formula

You may already be thinking ahead and realizing that structural formulas take up a lot of space and may be tedious to draw. **Condensed structural formulas** save space and time but still supply the necessary information contained in the structural formula. When you write the condensed structural formula, the C—H is understood. Therefore, if a carbon atom is joined to three hydrogen atoms, you can simply write CH_3. If the carbon atom has two hydrogen atoms, you would write CH_2. Using this convention, the first structural formula for butane could be written as:

$CH_3-CH_2-CH_2-CH_3$

or

$CH_3CH_2CH_2CH_3$

The second variation omits all lines, and assumes the carbons are all attached to each other and that the hydrogens are attached to the preceding carbon.

Condensing the second structural formula from above is slightly different. You can show the major chain of three carbons in the same manner you condensed the first formula, but the fourth carbon that branches off the main chain can be shown in two ways. The first option is to show the branch with the line representing the carbon-carbon bond:

$$\begin{array}{ccccc} & & CH_3 & & \\ & & | & & \\ CH_3 & - & CH_3 & - & CH_3 \end{array}$$

The second option is to condense the formula even further by placing the branches off of the main chain in brackets:

$CH_3CH(CH_3)CH_3$

Using this representation, the (CH_3) represents the branch off the main chain. The brackets are used to show that it is attached to the preceding carbon, as a branch. If it were a part of the main chain, it would be a CH_2, and not in brackets. This method of condensing the structural formula will probably be of most use to you, because drawing lines and ensuring they are in the correct positions can be rather time consuming. Keep in mind that as you learn this concept you may wish to draw the expanded formula first, to make sure there are no errors.

Finally, when there are many carbons in a hydrocarbon, a third type of notation is used:

Decane:

$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

$CH_3(CH_2)_8CH_3$

You can indicate repeating CH_2s by using brackets and a subscript to indicate the number of repeating units. Thus, the $(CH_2)_8$ means there are eight repeating CH_2s in a row, with no branches. However, note that in this course you will generally be working with parent chains up to only six carbons long.

Learning Activity 6.2

Properties of Normal Alkanes

A paraffin compound is an example of an alkane. Remember that alkanes are saturated hydrocarbons having the general formula C_nH_{2n+2}, (where C is a carbon atom, H is a hydrogen atom, and *n* is the number of carbons). The paraffins are major constituents of natural gas and petroleum. Paraffins containing fewer than five carbon atoms per molecule are usually gaseous at room temperature, while those having five to fifteen carbon atoms are usually liquids, and the straight chain paraffins having more than 15 carbon atoms per molecule are solids.

In this learning activity, you will plot a graph to illustrate the relationship between the number of carbons in the parent chain of an alkane and its melting and boiling point.

Properties of Paraffin Hydrocarbons		
Number of Carbons in Parent Chain	**Melting Point (°C)**	**Boiling Point (°C)**
1	–180	–160
2	–190	–90
3	–175	–40
4	–140	0
5	–140	40
6	–90	80
7	–95	100
8	–50	140
9	–45	150
10	–20	180
11	–20	195
12	0	220
13	–5	250
14	5	270
15	15	290

continued

Learning Activity 6.2 (continued)

1. Use the data table provided to complete the graph below. Plot *temperature* on the *y*-axis and *number of carbon atoms* on the *x*-axis. Your graph will show two curves, one for boiling point and one for melting point. To differentiate the two curves, use two different colours or draw one curve using a solid line and the other curve using a broken line.

2. What relationship exists among melting point (MP), boiling point (BP), and the number of carbons in the parent chain?

Lesson Summary

In this lesson, you added to your growing knowledge of hydrocarbons. You learned how the bonding characteristics of the carbon atom in hydrocarbons can result in single, double, or triple bonds. Then, you were introduced to the molecular structures of alkanes, alkenes, and alkynes. In the next lesson, you will learn how to name, draw, and construct structural models of both straight chain and branched chain alkanes.

Assignment 6.2

Aliphatic Hydrocarbons *(5 marks)*

1. If an alkane has

   a) 6 carbons, how many H atoms will it have? *(1 mark)*

   b) 20 H atoms, how many C atoms will it have? *(1 mark)*

   c) 25 C atoms, how many H atoms will it have? *(1 mark)*

   d) 12 H atoms, how many C atoms will it have? *(1 mark)*

2. Write the condensed structural formula for hexane (note that this is not the same thing as the molecular formula): $CH_3CH_2CH_2CH_2CH_2CH_3$. *(1 mark)*

NOTES

Lesson 3: Alkanes (2 hours)

Lesson Focus

SLO C11-5-05: Name, draw, and construct structural models of the first 10 alkanes.

Include: IUPAC nomenclature, structural formulas, condensed structural formulas, molecular formulas, general formula $C_nH_{(2n+2)}$

SLO C11-5-06: Name, draw, and construct structural models of the branched alkanes.

Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature

Lesson Introduction

Now that you have a general knowledge of hydrocarbons, you will study each individual classification of hydrocarbon. This lesson will focus on the first 10 alkanes, as well as some branched alkanes. You will find out how to name, draw, and construct structural models of the straight and branched alkanes.

You will be using your molecular model kit in this lesson. If you do not have access to this kit, then you can construct your molecules by using marshmallows and toothpicks.

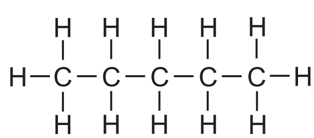
Naming Straight Chain Alkanes

In this lesson, you will learn to name organic compounds following the rules set out by the International Union of Pure and Applied Chemistry, or IUPAC. The rules involve choosing a main or parent chain and then using a series of prefixes and suffixes (or endings) to name the structure. In this course, you will only name hydrocarbons with a parent chain having up to 10 carbons.

First of all, *all alkanes* will end in *–ane*. Their prefixes describe the number of carbons in the parent chain. You will need to know these prefixes:

Number of Carbons	Prefix
1	meth–
2	eth–
3	prop–
4	but–
5	pent–
6	hex–
7	hept–
8	oct–
9	non–
10	dec–

Remember from the last lesson that alkanes can have both straight and branched chains. In the case of straight chain alkanes, every carbon is linked to two other carbons (except for the end, or terminal, carbons). For example, a straight chained alkane with five carbons would be called *pentane* and would look like the diagram below:



The table below gives a list of the 10 straight chained parent chains you will need to know, plus an 18-carbon chain for comparison.

Number of Carbons	Molecular Formula (C_nH_{2n+2})	Prefix
1	CH_4	methane
2	C_2H_6	ethane
3	C_3H_8	propane
4	C_4H_{10}	butane
5	C_5H_{12}	pentane
6	C_6H_{14}	hexane
7	C_7H_{16}	heptane
8	C_8H_{18}	octane
9	C_9H_{20}	nonane
10	$C_{10}H_{22}$	decane
18	$C_{18}H_{38}$	octadecane

Naming Branched Chain Alkanes

When naming branched chain alkanes, the rules are slightly more involved. First, an atom or group of atoms, called a **substituent**, can take the place of a hydrogen atom on a parent hydrocarbon molecule.

A hydrocarbon substituent is called an **alkyl group** and includes any branch that is a hydrocarbon with only single bonds. All alkyl groups use the prefixes for the number of carbons followed by the ending *–yl*. Below is a list of common alkyl groups you will use:

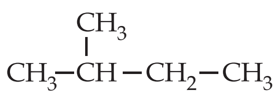
Formula	Name
$–CH_3$	methyl
$–C_2H_5$ or $–CH_2CH_3$	ethyl
$–C_3H_7$ or $–CH_2CH_2CH_3$	propyl

Basically, alkyl groups are alkanes that have had one terminal (or end) hydrogen removed. They are named by removing the *–ane* ending of the parent chain and replacing it with *–yl*. Remember that any alkane with one or more alkyl groups is automatically a branched chain alkane. In the case of branched alkanes, the terminal carbon is only linked to one other carbon.

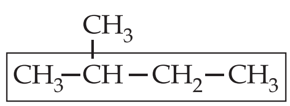
A Sample Naming Problem

Example 1

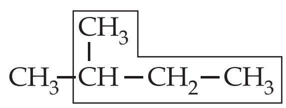
Name the following branched-chain alkane.



Step 1: *Find the longest continuous chain of carbons in the molecule – this is the parent chain.* If you are given the condensed structural formula, expand the formula to show all the carbon-carbon bonds.

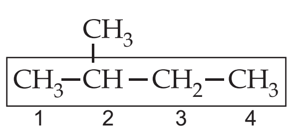


or

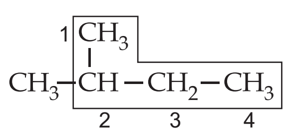


In this example, the parent chain is *butane*.

Step 2: *If there are branches off the parent chain (alkyl groups), number the carbon atoms in the parent chain so that the branches are on the lowest numbered carbons.*



or



In this example, the branch is on carbon #2, if numbered from left to right. If you numbered from right to left, the branch would be on carbon #3. Since the second numbering results in a higher number, you would choose to number from left to right.

Step 3: *Name and identify the position of the alkyl group.*

Each alkyl group is named according to the number of carbons, and the number of the carbon on the main chain to which it is attached. The name of each alkyl group is attached to the front of the parent chain with a hyphen separating the number of its location in the chain from the name of the alkyl group. **If there is more than one alkyl group, they are placed in alphabetical order.** Also, when there are more than one of the same alkyl group present in a hydrocarbon, you must use the prefixes *di*– (2), *tri*– (3), *tetra*– (4), *penta*– (5), et cetera. Each multiple alkyl group receives a number to indicate to which carbon it is attached, separated by commas.

In this example, there is only one methyl group on the #2 carbon.

Therefore, the name would be *2–methylbutane*.

Note that there are no spaces between any numbers, hyphens, or names.

Example 2

Name the following alkane:

$$CH_3CH(CH_3)CH_2CH(CH_3)CH_3$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

$$\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & - & CH & - & CH_2 & - & CH & - & CH_3 \end{array}$$

Step 2: *Choose the longest continuous carbon chain as the parent chain.*

$$\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ \end{array}$$
$$\boxed{CH_3-CH-CH_2-CH-CH_3}$$

Step 3: *Number the carbons to give any branches the lowest possible numbers.*

$$\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ \end{array}$$
$$\boxed{CH_3-CH-CH_2-CH-CH_3}$$
$$\begin{array}{ccccc} 1 & 2 & 3 & 4 & 5 \end{array}$$

In this case, numbering from either end gives the same result.

The parent chain is five carbons long, so it is *pentane*.

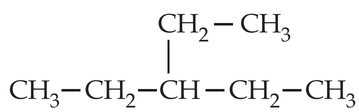
Step 4: *Identify the alkyl groups and name the compound.*

There is a methyl group on the #2 carbon, and another methyl group on the #4 carbon. Since there are two methyl groups, we must use the *di–* prefix. The name of the alkyl groups is *2,4–dimethyl*.

The name of the compound is *2,4–dimethylpentane*.

Example 3

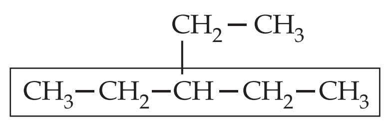
Name the following alkane:



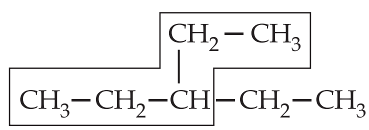
Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

The compound is already in expanded formula, so go to the next step.

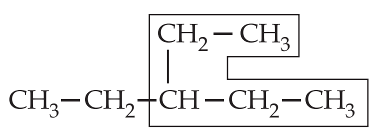
Step 2: *Choose the longest continuous carbon chain as the parent chain.*



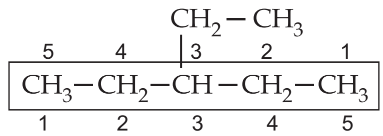
or



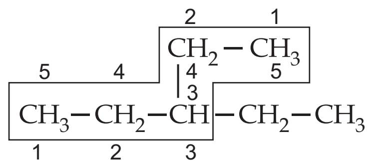
or



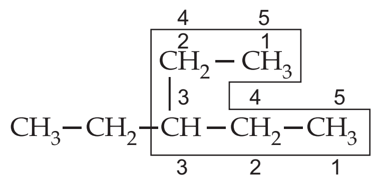
Step 3: *Number the carbons to give any branches the lowest possible numbers.*



or



or



In this case, numbering from either end gives the same result, regardless of which representation you draw.

The parent chain is five carbons long, so it is *pentane*.

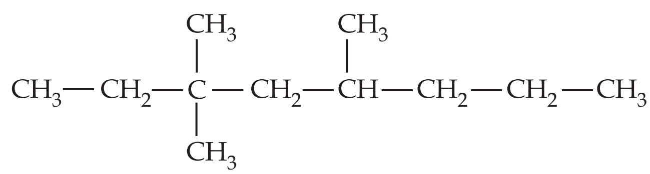
Step 4: *Identify the alkyl groups and name the compound.*

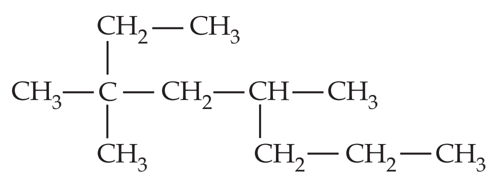
Since there is an ethyl group on the #3 carbon, the name of the alkyl group is *3-ethyl*.

The name of the compound is *3–ethylpentane*.

Extra Challenge

Here is an extra-challenging example for you to try: Give the IUPAC name for the following alkane. Note that while there are two representations shown, both are the same alkane.

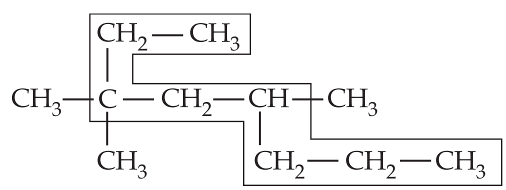




Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

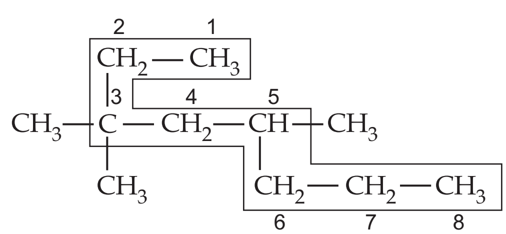
The structural formula is already expanded, so you can proceed to Step 2.

Step 2: *Choose the longest continuous carbon chain as the parent chain.*



Note that the parent chain does not need to be a straight line. You can wind around corners to find the longest chain.

Step 3: *Number the carbons to give any branches the lowest possible numbers.*



The parent chain is eight carbons long, or *octane*.

Step 4: *Identify the alkyl groups and name the compound.*

There are three methyl groups, two on the #3 carbon, and one on the #5 carbon. Therefore, you would name the alkyl group *3,3,5-trimethyl*.

> The name of the compound is *3,3,5–trimethyloctane*.

Drawing the Structural Formula Using the Name

You can also construct the structural formulas for organic compounds based on IUPAC names. When drawing a compound from its name, you would use the same principles as writing its name:

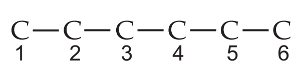
1. *Use the root word of the parent hydrocarbon (ending in –ane) to write out a parent (carbon) chain of the appropriate length.*
2. *Identify any alkyl groups and attach them at the proper positions to the parent chain.*
3. *Add hydrogens as needed so that every carbon atom has four bonds.*
4. *If necessary, condense the structural formula.*

Example 1

Draw the structural formula for 2-methylhexane.

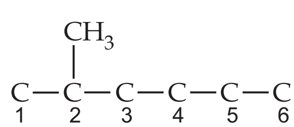
Step 1: *Use the root word of the parent hydrocarbon (ending in –ane) to write out a parent chain of the appropriate length.*

The parent chain is *hexane*, so it will be six carbons long.



Step 2: *Identify any alkyl groups and attach them at the proper positions to the parent chain.*

In this case, a single methyl group goes on carbon #2.

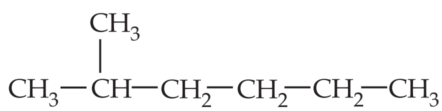


Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

Carbons 1 and 6 have only 1 bond, so they need 3 Hs.

Carbons 3, 4, and 5 each have 2 bonds, so they each need 2 Hs.

Carbon 2 has 3 bonds, so it needs only 1 H.



Step 4: *If necessary, condense the structural formula.*

Put the methyl group in brackets following the #2 carbon.

$CH_3CH(CH_3)CH_2CH_2CH_2CH_3$

Example 2

Draw the structural formula for 2,2-dimethylbutane.

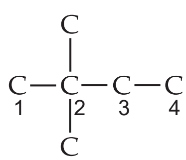
Step 1: *Use the root word of the parent hydrocarbon (ending in –ane) to write out a parent chain of the appropriate length.*

The parent chain is *butane*, so it will be four carbons long.

$C—C—C—C$

Step 2: *Identify any alkyl groups and attach them at the proper positions to the parent chain.*

In this case, there are two methyl groups on carbon #2.

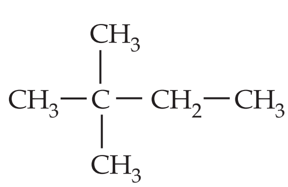


Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

Carbons 1 and 4 have only 1 bond, so they need 3 Hs.

Carbon 3 has 2 bonds, so it needs 2 Hs.

Carbon 2 has 4 bonds, so it doesn't need any Hs.



Step 4: *If necessary, condense the structural formula.*

Put the methyl groups in brackets following the #2 carbon.

$CH_3C(CH_3)_2CH_2CH_3$

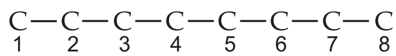
Extra Challenge

Here is a challenging example! Try it, but note that you will not be asked to draw this on assignments or the final exam.

Draw the structural formula for 3-ethyl-2,7-dimethyl-4-propyloctane.

Step 1: *Use the root word of the parent hydrocarbon (ending in –ane) to write out a parent chain of the appropriate length.*

The parent chain is *octane*, and it therefore has eight carbons.



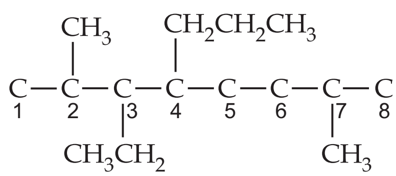
Step 2: *Identify any alkyl groups and attach them at the proper positions to the parent chain.*

In this example, there are several alkyl groups:

An ethyl (2 carbons) on carbon #3.

Two methyls on carbons #2 and 7.

A propyl (3 carbons) on carbon #4.



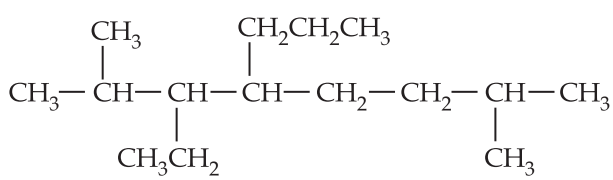
Note: The branches can be placed above or below the parent chain. It is easier to place all the branches if you spread them out, placing some above and some below the parent chain.

Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

Carbons 1 and 8 only have 1 bond, so they need 3 Hs each.

Carbons 5 and 6 have 2 bonds each and need 2 Hs each.

Carbons 2, 3, 4, and 7 have 3 bonds and only need 1 H each.



Step 4: *If necessary, condense the structural formula.*

$CH_3CH(CH_3)CH(CH_2CH_3)CH(CH_2CH_2CH_3)CH_2CH_2CH(CH_3)CH_3$

or

$CH_3CH(CH_3)CH(C_2H_5)CH(C_3H_7)CH_2CH_2CH(CH_3)CH_3$

Constructing Structural Models of Alkanes

So far in this lesson you have learned about naming and drawing alkanes based on their molecular formula and their two-dimensional structural form. While this is a convenient method to show the arrangement of carbons and hydrogens in an alkane, it does not allow you to see how the atoms are arranged in space. In order to better visualize this three-dimensional aspect of a compound's arrangement, you will construct models of alkanes using the molecular model kit that you purchased with this course. If you did not purchase this kit, then you can construct the molecules by using toothpicks and marshmallows.

Follow these steps to begin the building process:

1. *Determine the appropriate length of the parent (carbon) chain.* You may already know this if you have drawn the structural formula of the molecule, or by analyzing the name of the molecule.
   - If you are using a model kit (sometimes referred to as a ball-and-stick model kit) you will need *one black ball for every carbon*, joined together by the sticks provided.
   - If you are using toothpicks and marshmallows, you will need one marshmallow for every carbon, joined together by a toothpick. It would be helpful if you could get coloured marshmallows, so that you could use *one coloured marshmallow for each carbon* (the same colour for all) and white ones for hydrogen.
2. *Identify any alkyl groups, and then attach them at the appropriate positions on the parent chain.*
3. *Add hydrogens (white balls or white marshmallows) as needed so that every carbon atom has four bonds.*

Example 1

Construct the 3-D molecular model to represent ethane.

Step 1: *Determine the appropriate length of the parent (carbon) chain.*

Ethane, the simplest alkane, has a parent chain of only two carbons.

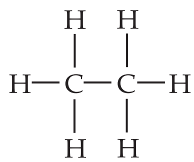
- Take two black balls and join them together with a stick, or
- Take two coloured marshmallows and join them together with a toothpick

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions on the parent chain.*

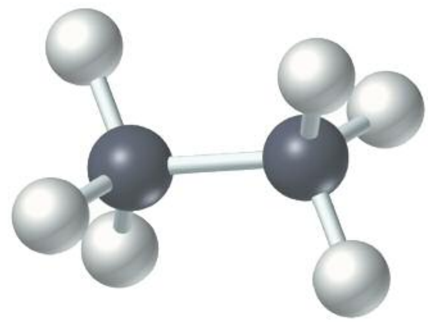
In this example there are no alkyl groups, so proceed to the next step.

Step 3: *Add hydrogens (white balls or white marshmallows) as needed so that every carbon atom has four bonds.*

Based on the hand-drawn, 2-D structural formula, you may know already that each carbon can bond to three hydrogens. You might also infer it based on the fact that every black ball has four holes, one for each bond that carbon can form. Take six hydrogens (white balls) and join them to each carbon with a stick. The 3-D model of ethane will look like this:



Ethane, C_2H_6



Example 2

Construct a 3-D molecular model to represent 2,4-dimethlypentane.

Step 1: *Determine the appropriate length of the parent (carbon) chain.*

Pentane has a parent chain of five carbons.

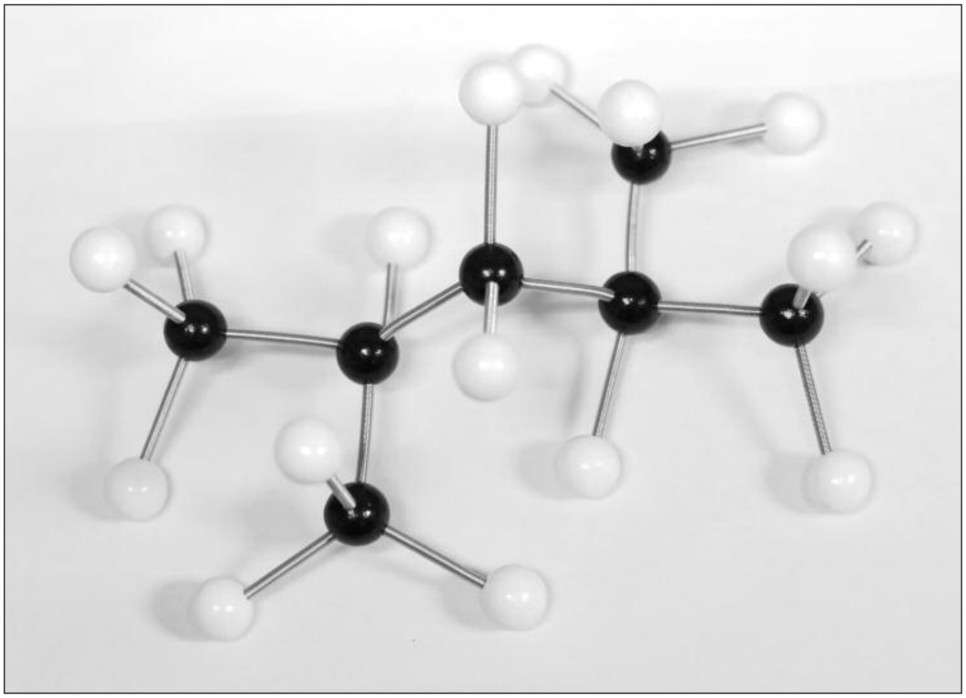
- Take five black balls and join them together with a stick, or
- Take five coloured marshmallows and join them together with a toothpick

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions on the parent chain.*

In this example, there are two methyl groups: one located on carbon #2, and another one on carbon #4. Using two more sticks (or toothpicks), join another black ball (or coloured marshmallow) to each of carbons 2 and 4.

Step 3: *Add hydrogens (white balls or white marshmallows) as needed so that every carbon atom has four bonds.*

To complete this 3-D structure, you will need to use a total of 16 white balls (or marshmallows). Join each hydrogen to a carbon using another stick (or toothpick). The 3-D model of 2,4-dimethlypentane will look like this:



Learning Activity 6.3

Working with Alkanes

1. How long can an alkyl group be? Draw and name the first three alkyl groups.
2. Name and draw the simplest straight chain alkane.
3. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it. (Hint: Draw the structures)
   a) 5–methylhexane
   b) 2–ethyl-2–methylpropane
   c) 1–methylbutane
   d) 3–ethylbutane
   e) 4–methyloctane
   f) 2,3–trimethylheptane
   g) 1,4–dimethylpentane
   h) 1,3,5–trimethylpentane
4. Name each of the following:
   a) $CH_3CH_2CH_3$
   b) $CH_3CH(CH_3)CH_2CH_2CH_3$
   c) $CH_3CH(CH_3)CH_2CH(CH_2CH_3)CH_2CH_2CH_3$
   d)

$$\begin{array}{ccc} & CH_3 & \\ & | & \\ CH_3 & C & CH_3 \\ & | & \\ & CH_3 & \end{array}$$

   e)

$$\begin{array}{cccccc} & CH_3 & & CH_3 & \\ & | & & | & \\ CH_3 & C & CH_2 & CH & CH_3 \\ & | & & & \\ & CH_3 & & & \end{array}$$

   f) $CH_3CH_2C(CH_3)_2CH_2CH_3$

continued

Learning Activity 6.3 (continued)

5. Draw the structural formulas for each of the following, condensing when possible:
   a) butane
   b) octane
   c) 2,3-dimethylpentane
   d) 3-ethylpentane
   e) 3-methylhexane
   f) 2,3-dimethylbutane
   g) 3-ethyl-2-methylpentane
   h) 3,3-diethylheptane
   i) 2,3,5-trimethylhexane
6. Construct 3-D models for the following alkanes:
   a) 2,2-dimethlypentane
   b) 2,3,3-trimethylpentane
   c) methane

Lesson Summary

In this lesson, you learned how to name, draw, and construct structural models of straight- and branched-chain alkanes. When naming alkanes, find the longest continuous chain of carbons and name it with the prefix representing the number of carbons and add the ending *–ane*. Number the carbons to give the alkyl branches the lowest numbers, and then name the branches according to the number of carbon atoms in the branch.

When drawing the structural formula for alkanes, draw the carbons in the parent chain, adding in the alkyl groups. Fill in hydrogen atoms until all of the carbons have four bonds. In the next lesson, you will discover what an isomer is, and learn to name and draw some isomers for alkanes.

Assignment 6.3

Naming and Drawing Alkanes *(15 marks)*

1. Name each of the following: *(7 marks)*

a) C_7H_{16}

____

b) $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

____

c)
$$\begin{array}{l} \quad\quad\quad\; CH_2CH_3 \\ \quad\quad\quad\; | \\ CH_3CH_2CHCH_2CH_2CH_2CH_3 \end{array}$$

____

d) $CH_3CH_2C(CH_3)_2CH_2CH_2CH_3$

____

e)
$$\begin{array}{ccccc} & CH_3 & & CH_3 & \\ & | & & | & \\ CH_3 & C & -CH_2- & C & CH_3 \\ & | & & | & \\ & CH_3 & & CH_3 & \end{array}$$

____

f) $CH_3CH_2CH(C_2H_5)CH_2CH_3$

____

g)
$$\begin{array}{ccccccccccc} & & & & & & CH_2 & - & CH_3 & & \\ & & & & & & | & & & & \\ CH_3 & - & CH_2 & - & CH & - & CH & - & CH_2 & - & CH_3 \\ & & & & | & & & & & & \\ & & & & CH_3 & & & & & & \end{array}$$

continued

Assignment 6.3: Naming and Drawing Alkanes (continued)

2. Draw the structural formulas for each of the following, condensing where possible. *(6 marks)*
   a) ethane

   b) decane

   c) 3-ethylhexane

   d) 2,2,3,4-tetramethylpentane

   e) 3-ethyl-2,2-dimethylhexane

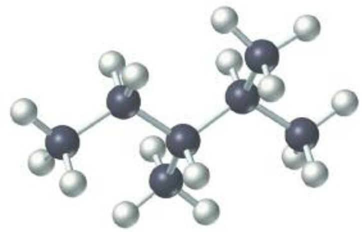
   f) 2,3-dimethylbutane

continued

Assignment 6.3: Naming and Drawing Alkanes (continued)

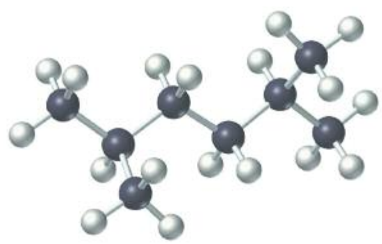
3. Name the following alkanes, based on the 3-D structural models shown below. *(2 marks)*

   a)



   ____

   b)



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NOTES

Lesson 4: Isomers of Alkanes (1 hour)

Lesson Focus

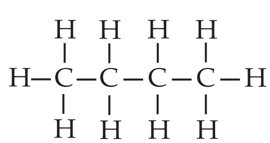
SLO C11-5-07: Name, draw, and construct structural models of isomers for alkanes up to six-carbon atoms.
Include: condensed structural formulas

Lesson Introduction

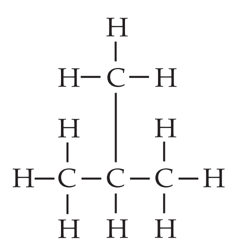
In the previous lesson, you learned general guidelines to help you draw and name branched and straight chain alkanes. In this lesson, you will go one step further and learn about isomers, eventually naming and drawing isomers for alkanes up to six carbon atoms long.

What is a Structural Isomer?

You may recall a previous example where two four carbon alkanes were compared:



butane



2-methylpropane

If you count the number of carbons and hydrogens in each compound, you will discover that both of these alkanes have the same molecular formula, C_4H_{10}. No doubt you have noticed that these two compounds have a different structural formula. Molecules, like the two above, that have the same molecular formula but different structural formulas are called structural isomers. Therefore, butane and 2-methylpropane are structural isomers.

Structural isomers will have the same molar mass, but will have different physical and chemical properties. The more carbon atoms a compound contains, the greater the number of structural isomers that can exist. For example:

Number of Carbon Atoms	Number of Isomers
4	2
5	3
6	5
8	18
10	75
20	3,660,319

Drawing Structural Isomers

It is easiest to work with a molecular formula, so if you are given a name, your first step should be to write the molecular formula. When drawing structural isomers, begin by drawing the straight carbon chain(s) without the hydrogen atoms, and then rearrange the carbon chain to find new patterns or arrangements. You should also try naming each arrangement to ensure they are actually different from one another. Complete each isomer by filling in the hydrogen atoms to give each carbon four bonds.

Example 1

Draw and name two structural isomers of pentane, not including pentane.

Step 1: *Write the molecular formula and draw a straight carbon chain.*

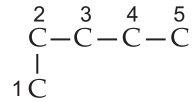
Pentane is a five carbon alkane, so its molecular formula is C_5H_{12}.

$CH_3–CH_2–CH_2–CH_2–CH_3$

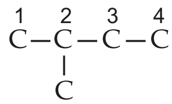
Step 2: *Draw different arrangements of the original carbon chain.*

There are many options. You can start by taking one carbon off the end and making it a branch. Start by just drawing the carbons, as you can add the hydrogens later. Then name the structure you have drawn to make sure you have created an isomer.

Note that not every arrangement will yield an isomer. For example, if you put a branch on the end, as in

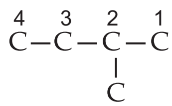


you have not created an isomer. Why not? Since the longest continuous carbon chain has five carbons, this structure is still the same as the original pentane. You attach the branches to internal carbons, not carbons on the end of the chain (also called the terminal carbons).

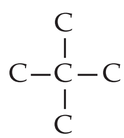


2-methylbutane

You could try putting the branch on the other end, like this:

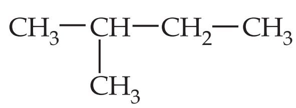


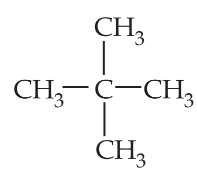
This makes a conformation of the original structure, but *not* another isomer, as this structure is still 2-methylbutane. Since you cannot make any more new structures with a single branch, try taking another carbon off the end and making a second branch.



2,2–dimethylpropane

After this process, you have determined that these are the only two structural isomers of pentane:





2–methylbutane   2,2–dimethylpropane

Constructing Structural Isomers

When constructing structural isomers, begin by building the straight carbon chain(s) without the hydrogen atoms, and then rearrange the carbon chain to find new patterns or arrangements. Again, try naming each arrangement to ensure that they are actually different from one another. Complete each isomer by adding the appropriate number of hydrogen atoms so that each carbon has four bonds.

You may want to refer to these steps when building each isomer:

1. *Determine the appropriate length of the parent (carbon) chain.* Remember that you will need one black ball or one coloured marshmallow for every carbon, joined together by sticks or toothpicks.
2. *Add any branches or alkyl groups to the proper positions on the parent chain.*
3. *Add hydrogens as needed so that every carbon atom has four bonds*, using either white balls or marshmallows.

Example 2

Draw and name two structural isomers of butane.

Step 1: *Write the molecular formula and draw a straight carbon chain.*

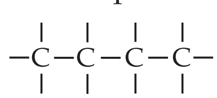
Butane is a four carbon alkane, so its molecular formula is C_4H_{10}, and its condensed structural formula is:

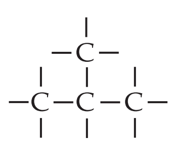
$CH_3–CH_2–CH_2–CH_3$

The parent chain is four carbons long. Connect four black balls or coloured marshmallows together. Next, try to find two different arrangements of four carbons.

Step 2: *Find new and different arrangements of the carbons.*

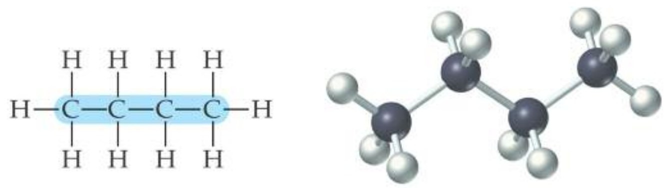
Since butane has only two internal carbons, you can probably predict that there will not be many isomers that you can create. If you remove a terminal carbon and add it to the middle of your now three carbon parent chain, you will create the following isomers:





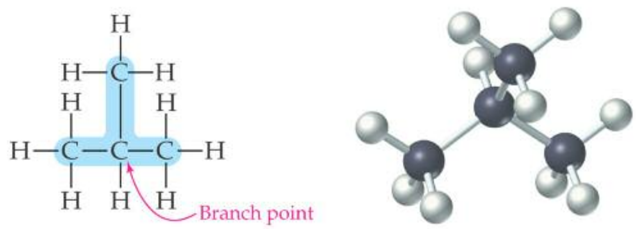
Step 3: *Complete each isomer by adding the appropriate number of hydrogen atoms to give each carbon four bonds.*

Now the two isomers would be drawn and constructed this way:



Straight chain

butane



Branched chain

2–methylpropane

Therefore the two isomers of butane are *butane* and *2–methylpropane*.

Identifying Isomers

Example 1

3-ethyl-2-methylhexane is a structural isomer of which straight chained alkane?

Step 1: *Identify the parent chain.*

The parent chain is *hexane*, which has six carbons.

Step 2: *Identify each branch and the number of carbons.*

There is one ethyl group (2 carbons) and 1 methyl group (1 carbon).

Step 3: *Find the sum of all of the carbons, and name the corresponding alkane.*

6 carbons + 2 carbons + 1 carbon = 9 carbons

The alkane with nine carbons is *nonane*.

3-ethyl-2-methylhexane is a structural isomer of *nonane*.

Learning Activity 6.4

Identifying Isomers of Alkanes

1. Indicate which straight chain alkane is a structural isomer of each of the following:
   a) 3–propylheptane
   b) 3–methylpentane
   c) 2,2,3,3–tetramethylpentane
   d) 2–methylbutane
   e) 3–ethylhexane
   f) 2–methylhexane
   g) 2,2–dimethylpentane
2. Define "structural isomer." Next, name and draw two compounds that are structural isomers.

Lesson Summary

In this lesson, you practiced identifying and drawing the structural isomers of alkanes. In the next lesson, you will transform alkanes to alkenes and vice-versa.

You will also learn how to name and draw alkenes and branched alkenes.

Assignment 6.4

Alkane Isomers *(15 marks)*

1. Draw, construct, and name the five structural isomers of hexane.

   Take digital photographs of the models you construct. You will then need to either mail or email these photographs for assessment. If this is a problem, contact your tutor/marker.

NOTES

Lesson 5: Alkenes (2 hours)

Lesson Focus

SLO C11-5-08: Outline the transformation of alkanes to alkenes and vice-versa.
Include: dehydrogenation/hydrogenation, molecular models

SLO C11-5-09: Name, draw, and construct molecular models of alkenes and branched alkenes.
Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature, structured formulas, condensed structural formulas, molecular formulas, general formula C_nH_{2n}

SLO C11-5-10: Differentiate between saturated and unsaturated hydrocarbons.

Lesson Introduction

In the previous lesson, you found out how to draw the isomers of alkanes up to six carbons. Now that you have practiced naming and drawing alkanes, you will move on to the next type of hydrocarbon, called an alkene. In this lesson, you will discover how to transform an alkane to an alkene. You will also learn how to name, draw, and construct molecular models of alkenes and branched alkenes.

Saturated and Unsaturated Hydrocarbons

Saturated hydrocarbons have the most hydrogens that the given number of carbons can hold. Alkanes are saturated compounds, since the only bonds in alkanes are single bonds. Compounds that contain double and triple bonds are therefore unsaturated. An unsaturated hydrocarbon does not contain the maximum number of hydrogens. When carbons form double and triple bonds, it means there is less bonding that can occur with hydrogens. This is why the ratio of hydrogen atoms to carbon atoms is lower in an unsaturated compound as compared to a saturated compound. Note the following examples of unsaturated compounds.

A double carbon-carbon bond is shown by two lines, rather than one:

$-C=C-$

A triple carbon-carbon bond is shown with three lines, rather than one:

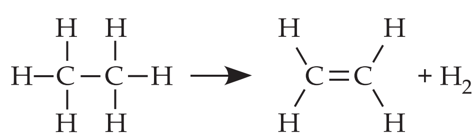
$-C\equiv C-$

Saturated hydrocarbons tend to have higher melting points and boiling points. Oils that are largely saturated fats are usually solid at room temperature. Butter, cheese, and some meats are examples of foods that are high in saturated fats. These types of fats are not healthy in that they raise levels of blood cholesterol, which is a risk factor for heart disease and stroke. The saturated hydrocarbons, because of their higher melting points, are more likely to be solid at room temperature. You will often see food products advertised as low in saturated fats and high in polyunsaturated fats. Polyunsaturated means the fat has many double and/or triple bonds. Unsaturated fats are required by the body for cell membranes and hormone production, and are not considered as unhealthy for the heart as saturated fats. Examples of unsaturated fats include olive oil and some types of margarine.

Transforming Alkanes to Alkenes

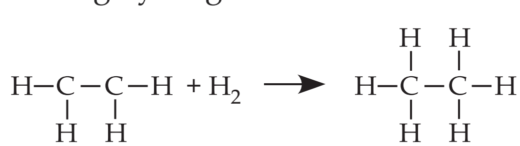
It is possible to transform an *alkane* into an *alkene* and an *alkene* into an *alkane*. These reactions are called dehydrogenation and hydrogenation.

1. Multiple carbon-carbon bonds can be formed by the removal of hydrogen. This process is called **dehydrogenation**. In other words, by removing hydrogen atoms, you can convert an alkane into an alkene. Here is an example that shows ethane converting to ethene:



Ethane **Ethene**

2. Alkenes are converted into alkanes by a reaction called **hydrogenation**, a common reaction in organic chemistry. In order for this reaction to proceed rapidly at normal temperatures, a catalyst such as nickel or platinum must be added. In the process below, ethene is transformed back into ethane by adding hydrogen:



This process is used to convert vegetable oils into margarine. “Trans” fats or “trans” fatty acids (TFAs) are formed when manufacturers hydrogenate unsaturated compounds with hydrogen to form saturated structures. Manufacturers have found that hydrogenating vegetable oils has many economic benefits. The process extends shelf life, increases the flavour stability of the product, and produces a solid product. There are, however, some studies that have suggested that TFAs present a greater risk of coronary artery disease than saturated fatty acids.

Naming Alkenes

An alkene has at least one double carbon-carbon bond. The general molecular formula for an alkene is:

C_nH_{2n}

According to the IUPAC rules for naming organic compounds (revised in 2008), alkenes are named for their parent alkane chain with the suffix “*–ene*” and an infixed number indicating the position of the first double-bonded carbon in the chain. For example, $CH_2{=}CHCH_2CH_3$ is *but–1–ene*. According to the old IUPAC naming system, this would be called 1–butene, or simply *butene* (as it is not necessary to write the number 1).

There are, however, two exceptions to these naming rules. *Ethene* and *propene* do not require infixed numbers, since there is no ambiguity in their structures. In other words, the double bond must be positioned on the first carbon in both cases. As these are the only exceptions, for all other alkenes you must number the position of the double bond.

Multiple double bonds are identified as *–diene*, *–triene*, etc., with the prefix of the chain having added an extra “a” added to it. For example, $CH_2{=}CHCH{=}CH_2$ is *buta–1,3–diene*.

Steps to follow when naming alkenes using the new IUPAC system:

1. *Expand the condensed formula to show all carbon-carbon bonds.*
2. *Choose the longest continuous carbon chain that contains the double bond as the parent chain. Number the carbons, ensuring the first carbon containing the double bond has the lowest number. Name the parent chain (ending in –ene), infixing the number of the first carbon with the double bond separated with a hyphen.*

   Reminder: Ethene and propene do not require infixed numbers, as there is no ambiguity in their structures.
3. *If there are multiple double bonds, assign the appropriate prefix (-diene, -triene, etc.) for the length of the parent chain having an extra “a” added to it. If there are any alkyl groups, name them as you did for the alkanes.*

Example 1

Name the following alkene:

$$CH_2 = CHCH_2CH_2CH_3$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

This is already a straight chained alkene, so you don't need to expand it.

Step 2: *Choose the longest continuous carbon chain that contains the double bond as the parent chain. Number the carbons, ensuring the first carbon containing the double bond has the lowest number. Name the parent chain (ending in –ene), infixing the number of the first carbon with the double bond separated with a hyphen.*

$$\overset{1}{C}H_2=\overset{2}{C}H\overset{3}{C}H_2\overset{4}{C}H_2\overset{5}{C}H_2$$

The double bond begins on carbon #1.

The compound is *pent–1–ene* (according to the old IUPAC rules, it would be called *1–pentene* or simply *pentene*).

Example 2

Name the following alkene:

$$CH_3CH = CHCH_2CH_3$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

Again, this is a straight chained alkene, so you don't need to expand it.

Step 2: *Choose the longest continuous carbon chain that contains the double bond as the parent chain. Number the carbons, ensuring the first carbon containing the double bond has the lowest number. Name the parent chain (ending in –ene), infixing the number of the first carbon with the double bond separated with a hyphen.*

$$\overset{1}{C}H_3\overset{2}{C}H=\overset{3}{C}H\overset{4}{C}H_2\overset{5}{C}H_3$$

The double bond begins on carbon #2.

The compound is *pent–2–ene* (or *2–pentene* using the old system).

Example 3

Name the following alkene:

$$\begin{array}{l} CH_3CH{=}CHCHCH_3 \\ \qquad\qquad\quad\;\; | \\ \qquad\qquad\quad\; CH_3 \end{array}$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

This is a straight chained alkene, so you don't need to expand it.

Step 2: *Choose the longest continuous carbon chain that contains the double bond as the parent chain. Number the carbons, ensuring the first carbon containing the double bond has the lowest number. Name the parent chain (ending in –ene), infixing the number of the first carbon with the double bond separated with a hyphen.*

$$\begin{array}{l} \overset{1}{C}H_3\overset{2}{C}H{=}\overset{3}{C}H\overset{4}{C}H\overset{5}{C}H_3 \\ \qquad\qquad\quad\;\; | \\ \qquad\qquad\quad\; CH_3 \end{array}$$

The double bond begins on carbon #2.

The compound is *pent–2–ene*.

Step 3: *If there are multiple double bonds, assign the appropriate prefix (–diene, –triene, etc.) with the size prefix of the chain, adding an extra "a." If there are any alkyl groups, name them as you did for the alkanes.*

There is a methyl group on carbon #4.

The name of the compound is *4–methyl–pent–2–ene*. The old IUPAC system would call this *4–methyl–2–pentene*. Note, the lowest number goes to the double bond.

Example 4

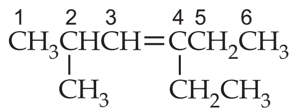
Name the following structure:

$$CH_3CH(CH_3)CH{=}C(C_2H_5)CH_2CH_3$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

$$\begin{array}{l} CH_3CHCH{=}CCH_2CH_3 \\ \qquad\;\, | \qquad\quad\; | \\ \qquad CH_3 \qquad CH_2CH_3 \end{array}$$

Step 2: *Choose the longest continuous carbon chain that contains the double bond as the parent chain. Number the carbons, ensuring the first carbon containing the double bond has the lowest number. Name the parent chain (ending in –ene), infixing the number of the first carbon with the double bond separated with a hyphen.*



You must number this six carbon alkene from right to left to identify the double bond by the lowest number, 3. Therefore the compound is *hex–3–ene*.

Step 3: *If there are multiple double bonds, assign the appropriate prefix (–diene, – triene, etc.) with the size prefix of the chain, adding an extra "a." If there are any alkyl groups, name them as you did for the alkanes.*

There is a methyl group on carbon #2 and an ethyl on carbon #4.

The name of the compound is *4–ethyl–2–methyl–hex–3–ene*. When naming the alkyl groups, list them alphabetically.

Drawing and Constructing Alkenes

The process of drawing and constructing alkenes is very similar to what you learned for alkanes.

Step 1: *Determine the appropriate length of the parent (carbon) chain. Draw the carbons for the parent chain, number the carbons, and insert the double bond.*

To build an alkene, you will need one black ball (or one coloured marshmallow) for every carbon, joined together by sticks or toothpicks. Where you are inserting the double bond, you will join two carbons with two sticks (occupying two of the four bonds that each of these carbons can form).

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

Keep in mind that every methyl group requires you to add one black ball or coloured marshmallow (for one carbon) while every ethyl group requires two black balls or coloured marshmallows (for two carbons).

Step 3: *Add hydrogens as needed so that every carbon atom has four bonds. Don't forget that the double bond counts as two bonds!*

On your model, add hydrogens as needed so that every carbon has formed four bonds, using either white balls or marshmallows. The carbons that have formed the double bond can only be joined to two hydrogens each.

Example 1

Draw the structural formula for but-1-ene (formerly called 1-butene).

Step 1: *Determine the appropriate length of the parent (carbon) chain. Draw the carbons for the parent chain, number the carbons, and insert the double bond.*

The parent chain is butene, therefore having four carbons. The double bond goes on the first carbon (that is, between carbons 1 and 2).

$$\overset{1}{C}=\overset{2}{C}-\overset{3}{C}-\overset{4}{C}$$

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

In this example, there are no alkyl groups.

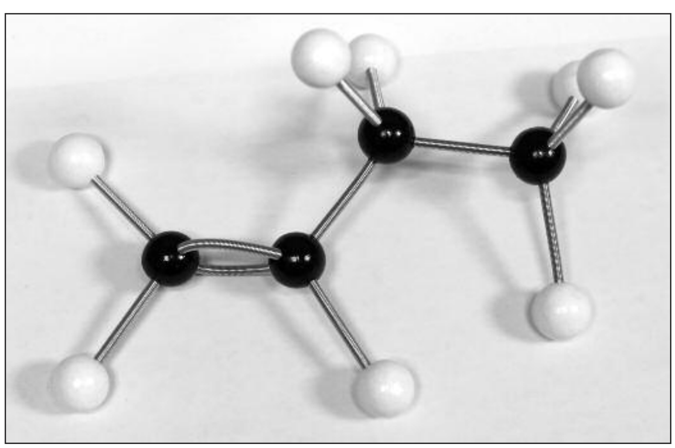
Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

$CH_2=CH-CH_2-CH_3$

If necessary, you can condense the structural formula for assignments. In this case, the formula is sufficiently condensed. Therefore, the structural formula of but-1-ene is:

$CH_2=CH-CH_2-CH_3$

The 3-D model would look like this:



Example 2

Draw the structural formula for 2–methylhex–2–ene.

Step 1: *Determine the appropriate length of the parent (carbon) chain. Draw the carbons for the parent chain, number the carbons, and insert the double bond.*

The parent chain is hexene, therefore having six carbons. The position of the double bond is on the second carbon.

$$\begin{array}{ccccccccccc} 1 & & 2 & & 3 & & 4 & & 5 & & 6 \\ C & - & C & = & C & - & C & - & C & - & C \end{array}$$

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

There is one methyl group on carbon #2.

$$\begin{array}{ccccccccccc} 1 & & 2 & & 3 & & 4 & & 5 & & 6 \\ C & - & C & = & C & - & C & - & C & - & C \\ & & | & & & & & & & & \\ & & CH_3 & & & & & & & & \end{array}$$

Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

$$\begin{array}{ccccccccccc} CH_3 & - & C & = & CH & - & CH_2 & - & CH_2 & - & CH_3 \\ & & | & & & & & & & & \\ & & CH_3 & & & & & & & & \end{array}$$

Since carbon #3 has a double bond and one carbon-carbon single bond, it only needs one more hydrogen.

Step 4: *If necessary, condense the structural formula.*

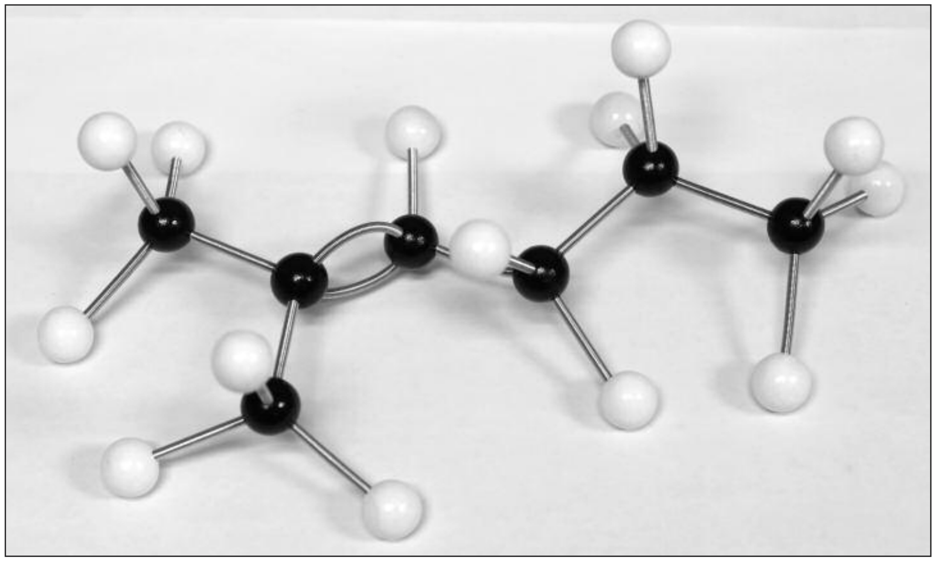
$CH_3-C(CH_3)=CH-CH_2-CH_2-CH_3$

or

$CH_3C(CH_3)=CH(CH_2)_2CH_3$

Notice that *you must still show the double bond* in the condensed structural formula.

The corresponding ball-and-stick model would look like this:



Common Alkenes

Ethene, the simplest alkene, is the gas released by fruit that promotes ripening. It is usually released by the maturing seeds of the fruit. Fruit producers will often place crates of fruit into vaults and expose them to ethene gas to speed up ripening. You can try this at home to speed up the ripening of store-bought fruit by placing unripe fruit into a paper bag containing a ripe apple. The ethene gas released by the apple will speed up the ripening of the other fruit in the bag.

Propene is commonly called propylene. When many molecules of an organic molecule are joined together, it forms a molecule called a polymer. Polypropylene, a plastic made up of propene molecules, is used in packaging, car bumpers, and film.

Learning Activity 6.5

Identifying Alkenes

1. What is the difference between saturated and unsaturated organic compounds? Give an example of each.
2. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it.
   a) 2–methylbut–3–ene
   b) 2,2–dimethylpent–3–ene
   c) 1–butene
   d) 2,2–dimethylhex–1–ene
   e) 3–methylprop–1–ene
   f) hex–5–ene
3. Write the balanced equation, using structural formulas, for the following reactions:
   a) The hydrogenation of but–2–ene
   b) The dehydrogenation of propane
   c) 4–methylhex–2–ene + $1H_2$
4. Name the products of the chemical equations in question #3 above.

Lesson Summary

In this lesson, you learned how alkanes can be converted to alkenes, as well as the rules for naming and drawing alkenes. In the next lesson, you will outline the transformation of alkenes to alkynes. You will also practice naming and drawing alkynes and branched alkynes.

Assignment 6.5

Naming and Drawing Alkenes *(16 marks)*

1. Draw the structural formulas for the following alkenes. *(7 marks)*
   a) prop–1–ene

   b) 4–ethylpent-2–ene

   c) 4,4–dimethylpent-2–ene

   d) 2–methylbut–1–ene

   e) 2,3,5–trimethylhex–2–ene

continued

Assignment 6.5: Naming and Drawing Alkenes (continued)

f) 2–methylbut–2–ene

g) hex–2–ene

2. Name the following compounds. *(6 marks)*

a) $CH_3CH{=}CHCH_3$

____

b) $$\mathrm{CH_3\underset{\underset{\displaystyle CH_3}{|}}{C}HCH_2CH{=}CHCH_3}$$

____

c) $H_2C{=}CH_2$

____

d) $$\mathrm{CH_2{=}CH\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}{-}CH_2\overset{\overset{\displaystyle CH_2CH_3}{|}}{C}HCH_2CH_3}$$

____

continued

Assignment 6.5: Naming and Drawing Alkenes (continued)

e) $$\begin{array}{l} CH_2{=}CHCH_2\overset{}{C}HCH_3 \\ \qquad\qquad\quad\;\; | \\ \qquad\qquad\quad\; CH_3 \end{array}$$

____

f) $$\begin{array}{l} CH_3CHCH_2CH_2CH{=}CH_2 \\ \qquad\;\; | \\ \qquad\; CH_2CH_3 \end{array}$$

____

3. Use your molecular model kit to construct models of the following alkenes. Then take a photograph (with either a digital camera or a film camera) of these models and mail or email them once you have completed this module. *(3 marks)*

a) 2,3,5–trimethylhex–2–ene

b) 2–methylbut–2–ene

continued

Assignment 6.5: Naming and Drawing Alkenes (continued)

c) hex–2–ene

Lesson 6: Alkynes (2 hours)

Lesson Focus

SLO C11-5-11: Outline the transformation of alkenes to alkynes and vice-versa.
Include: dehydrogenation/hydrogenation, molecular models

SLO C11-5-12: Name, draw, and construct molecular models of alkynes and branched alkynes.
Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature, structured formulas, condensed structural formulas, molecular formulas, general formula C_nH_{2n-2}

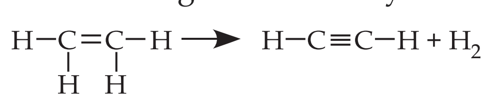
Lesson Introduction

In the last lesson, you focused on the study of alkenes, including the rules for naming and drawing alkenes. In this lesson, you will continue your study of hydrocarbons by examining the alkynes. You will outline the transformation of alkenes to alkynes, and will learn the rules for naming and drawing alkynes.

Alkynes from Alkenes

The process involved in transforming alkenes to alkynes is similar to what you saw in the previous lesson. The reactions that convert between alkanes and alkenes are the same ones that convert between alkenes and alkynes—dehydrogenation and hydrogenation.

Triple carbon-carbon bonds can be formed by the further dehydrogenation of alkenes. In other words, by removing hydrogen atoms from a double bond, you create a triple bond, thus converting an alkene to an alkyne. Here is an example converting ethene to ethyne:



Alkynes can be converted to alkenes by hydrogenation, a common reaction that you were introduced to in the last lesson. In order for this reaction to proceed rapidly at normal temperatures, a catalyst such as nickel or platinum must be added. In the process below, ethyne is transformed back into ethene by adding hydrogen:

$$H-C\equiv C-H + H_2 \longrightarrow H-\underset{\displaystyle H}{\underset{|}{C}}=\underset{\displaystyle H}{\underset{|}{C}}-H$$

Naming Alkynes

Alkynes are unsaturated hydrocarbons with one or more triple carbon-carbon bond. Since two more hydrogens, as compared to alkenes, are replaced by the extra bond, the general molecular formula for alkynes is:

$$C_nH_{2n\text{-}2}$$

Naming alkynes is similar to naming alkenes, except alkynes end in *–yne*. According to the IUPAC rules of naming (revised in 2008), alkynes are named for their parent alkane chain with the suffix "–yne" and an infixed number indicating the position of the first triple-bonded carbon in the chain. For example, $CH\equiv CCH_2CH_3$ is *but–1–yne*. According to the old IUPAC naming system, this would be called *1–butyne*.

There are, however, two exceptions to these naming rules. *Ethyne* and *propyne* do not require infixed numbers, since there is no ambiguity in the structures. In other words, the triple bond must be positioned on the first carbon in both cases. These are the only exceptions. For all other alkynes, you must number the position of the triple bond.

The Steps for Naming Alkynes

1. *Expand the condensed formula to show all carbon-carbon bonds.*
2. *Choose the longest continuous carbon chain that contains the triple bond as the parent chain. Number the carbons, ensuring the first carbon containing the triple bond has the lowest number. Name the parent chain (ending in –yne), infixing the number of the first carbon with the triple bond, separated with a hyphen.*
3. *If there are any alkyl groups, name them as you did for the alkanes.*

Example 1

Name the following alkyne:

$$CH_3-C\equiv C-CH_2-CH_3$$

Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

This is already a straight chained alkyne, so you don't need to expand it.

Step 2: *Choose the longest continuous carbon chain that contains the triple bond as the parent chain. Number the carbons, ensuring the first carbon containing the triple bond has the lowest number. Name the parent chain (ending in –yne), infixing the number of the first carbon with the triple bond separated with a hyphen.*

$$\overset{1}{C}H_3-\overset{2}{C}\equiv\overset{3}{C}-\overset{4}{C}H_2-\overset{5}{C}H_3$$

The triple bond begins on carbon #2. The compound is *pent-2-yne*.

Step 3: *If there are any alkyl groups, name them as you did for the alkanes.*

There are no alkyl branches. Therefore, the name of the compound is simply *pent–2–yne*.

Example 2

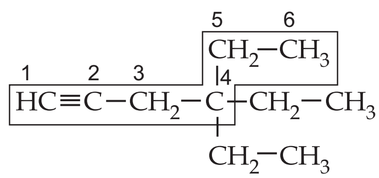
Name the following alkyne:

$$HC\equiv CCH_2C(C_2H_5)_2CH_2CH_3$$

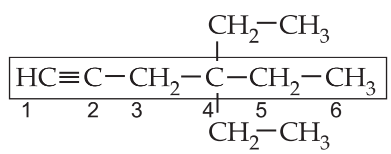
Step 1: *Expand the condensed formula to show all carbon-carbon bonds.*

$$\begin{array}{c}
\qquad\qquad\qquad\quad CH_3CH_2 \\
\qquad\qquad\quad | \\
HC\equiv C\text{–}CH_2-C-CH_2-CH_3 \\
\qquad\qquad\quad | \\
\qquad\qquad\qquad\quad CH_3CH_2
\end{array}$$

Step 2: *Choose the longest continuous carbon chain that contains the triple bond as the parent chain. Number the carbons, ensuring the first carbon containing the triple bond has the lowest number. Name the parent chain (ending in –yne), infixing the number of the first carbon with the triple bond separated with a hyphen.*



or



The triple bond is located on carbon #1 and the parent chain is six carbons long. Therefore, the compound is *hex–1–yne*.

Step 3: *If there are any alkyl groups, name them as you did for the alkanes.*

There are two ethyl groups on the #4 carbon. The name of the compound is *4,4–diethylhex–1–yne*.

Drawing Alkynes

The rules for drawing alkynes are similar to those used for drawing alkenes. When drawing alkynes:

1. *Determine the appropriate length of the parent chain. Draw the carbons for the parent chain, number the carbons, and insert the triple bond.*
2. *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*
3. *Add hydrogens as needed so that every carbon atom has four bonds. Remember that the triple bond counts as three bonds!*
4. *If necessary, condense the structural formula.* In other words, put the alkyl groups in brackets following the carbon to which it is attached.

Example 1

Draw the structural formula for pent–1–yne.

Step 1: *Determine the appropriate length of the parent chain. Draw the carbons for the parent chain, number the carbons, and insert the triple bond.*

The parent chain is pentyne, therefore having five carbons. The triple bond goes on the first carbon (between carbons 1 and 2).

$$\overset{1}{C}\equiv\overset{2}{C}-\overset{3}{C}-\overset{4}{C}-\overset{5}{C}$$

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

There are no alkyl groups.

Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

$$CH\equiv C-CH_2-CH_2-CH_3$$

Step 4: *If necessary, condense the structural formula.*

This is sufficiently condensed. The structural formula of pent-1-yne is

$$CH\equiv C-CH_2-CH_2-CH_3$$

Example 2

Draw the structural formula for 2,5–dimethylhex–3–yne.

Step 1: *Determine the appropriate length of the parent chain. Draw the carbons for the parent chain, number the carbons, and insert the triple bond.*

The parent chain is hexyne, therefore having six carbons. The triple bond goes on the third carbon (between carbons 3 and 4).

$$\overset{1}{C}-\overset{2}{C}-\overset{3}{C}\equiv\overset{4}{C}-\overset{5}{C}-\overset{6}{C}$$

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

There two methyl groups, one on carbon #2 and one on carbon #5.

$$\begin{array}{ccccccccccc} & & CH_3 & & & & & & CH_3 & & \\ & & | & & & & & & | & & \\ C & - & C & - & C & \equiv & C & - & C & - & C \end{array}$$

Step 3: *Add hydrogens as needed so that every carbon atom has four bonds.*

$$\begin{array}{ccccccccccc} & & CH_3 & & & & & & CH_3 & & \\ & & | & & & & & & | & & \\ CH_3 & - & CH & - & C & \equiv & C & - & CH & - & CH_3 \end{array}$$

Step 4: *If necessary, condense the structural formula.*

The condensed structural formula of 2,5-dimethylhex-3-yne would be:

$CH_3CH(CH_3)C \equiv CCH(CH_3)CH_3$

Constructing Structural Models of Alkynes

The process of drawing and constructing alkynes is very similar to what you learned for alkenes.

Step 1: *Determine the appropriate length of the parent (carbon) chain. Assemble the correct number of carbons for the parent chain, and insert the triple bond at the correct location.*

You will need one black ball (or one coloured marshmallow) for every carbon, joined together by sticks or toothpicks. Where you are inserting the triple bond, you will join two carbons with three sticks (occupying three of the four bonds that these carbons can form).

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

Keep in mind that every methyl group requires you to add one black ball or coloured marshmallow (for one carbon) while every ethyl group requires two black balls or coloured marshmallows (for two carbons).

Step 3: *Add hydrogens (white balls or marshmallows) as needed, so that every carbon has four bonds.* Don't forget that the triple bond counts as three bonds!

Example 1

Draw and construct the structure for ethyne.

Step 1: *Determine the appropriate length of the parent (carbon) chain. Assemble the correct number of carbons for the parent chain, and insert the triple bond at the correct location.*

In this case, the parent chain is made up of two carbons. You will need to join two black balls or coloured marshmallows together using three sticks or toothpicks.

Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

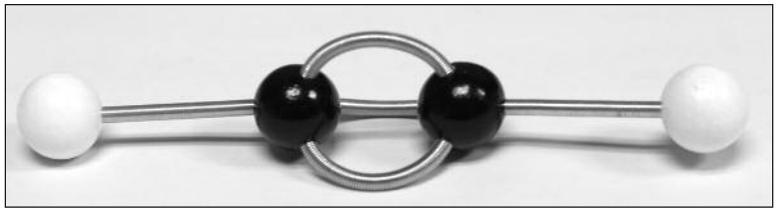
In this example, there are no alkyl groups.

Step 3: *Add hydrogens (white balls or marshmallows) as needed, so that every carbon has four bonds.*

In this case, each carbon can link to one hydrogen atom. The final model of ethyne should look like this:

$H-C\equiv C-H$

Acetylene, C_2H_2



Example 2

Draw and construct the structure for 4,4–dimethylpent–2–yne.

Step 1: *Determine the appropriate length of the parent (carbon) chain. Assemble the correct number of carbons for the parent chain, and insert the triple bond at the correct location.*

In this case, the parent chain (pentyne) is made up of five carbons, with the triple bond placed after the second carbon. You will need to join five black balls or coloured marshmallows together, using three sticks or toothpicks between the second and third carbons.

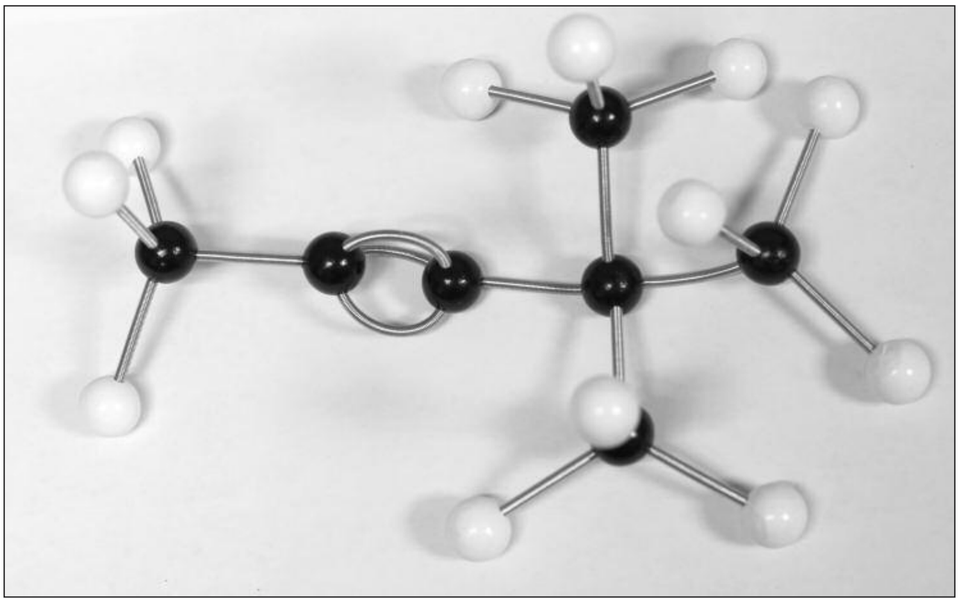
Step 2: *Identify any alkyl groups, and then attach them at the appropriate positions to the parent chain.*

In this example, there are two methyl groups, both located at carbon #4. Join two carbons (black balls or coloured marshmallows) to carbon #4.

Step 3: *Add hydrogens (white balls or marshmallows) as needed, so that every carbon has four bonds.*

The final model of 4,4–dimethylpent–2–yne should look like this:

$$\begin{array}{ccccccccc} & & & & & & CH_3 & & \\ & & & & & & | & & \\ CH_3 & - & C & \equiv & C & - & C & - & CH_3 \\ & & & & & & | & & \\ & & & & & & CH_3 & & \end{array}$$



Common Alkynes

Alkynes are not plentiful in nature. *Acetylene*, C_2H_2, is the common name for ethyne. It is the simplest of all alkynes and can be used as a starting molecule for the production of larger alkyne molecules. The manufacture of acetylene can be accomplished by the reaction of water with calcium carbide (CaC_2), which is manufactured by the reaction of calcium oxide and coal under high temperature. Acetylene is the fuel burned in oxyacetylene torches, which are commonly used in welding. Acetylene burns very clean and hot in the presence of pure oxygen.

Acetylene is also the organic starting material for the large–scale production of many important organic compounds including acetic acid (also known as vinegar), plastics, and synthetic rubber.

Learning Activity 6.6

Identifying Alkynes

1. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it.
   a) hex–4–yne
   b) 2,2–dimethyl–3–pentyne
   c) but–1–yne
   d) 2,2–dimethylhex–2–yne
   e) 3–methylprop–2–yne
2. Write the balanced equation, using structural formulas, for the following reactions:
   a) hydrogenation of but–2–yne
   b) dehydrogenation of prop–1–ene
   c) 4–methylhex–2–yne + $1H_2$
3. Name the products formed in question 2 above.

Lesson Summary

In this lesson, you learned that alkynes have at least one triple carbon-carbon bond, and have the molecular formula C_nH_{2n-2}. When naming alkynes, the parent chain is numbered to give the first carbon in the triple bond the lowest number and the compound name ends in *–yne*. In the next lesson, you will compare and contrast aromatic and aliphatic hydrocarbons, as well as look at some uses of aromatic hydrocarbons.

NOTES

Assignment 6.6

Naming and Drawing Alkynes *(15 marks)*

1. Draw the molecular structures for the following alkynes. *(6 marks)*
   a) prop–1–yne

   b) 5–methylhex–2–yne

   c) hex–3–yne

   d) 4,4–dimethylhex–2–yne

   e) 3–methylbut–1–yne

continued

f) 3–ethylpent–1–yne

2. Name the following alkynes. *(7 marks)*

a) $CH_3C\equiv CCH_3$

____

b) $CH_3\underset{\underset{CH_3}{|}}{C}HCH_2C\equiv CCH_3$

____

c) $HC\equiv CH$

____

d) $CH\equiv C\overset{\overset{CH_3}{|}}{C}HCH_2\overset{\overset{CH_3}{|}}{C}HCH_3$

____

e) $CH_3C\equiv CC(C_2H_5)_2CH_2CH_3$

____

f) $CH_3C\equiv CC(CH_3)_2CH_3$

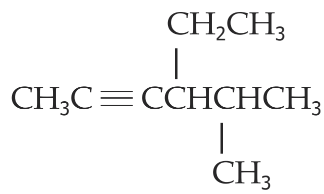
____

continued

Assignment 6.6: Naming and Drawing Alkynes (continued)

g) $CH_3C \equiv CCH(C_2H_5)CH(CH_3)CH_3$

or



3. Use your molecular model kit to construct models of the following alkynes. Then take a photograph (with either a digital camera or a film camera) of these models, and mail or email them once you have completed this module. *(2 marks)*

   a) 4,4–dimethylhex–1–yne

   b) pent–2–yne

NOTES

Lesson 7: Aromatic Hydrocarbons (1 hour)

Lesson Focus

SLO C11-5-13: Compare and contrast the aromatic and aliphatic hydrocarbons.
Include: molecular models, condensed structural formulas

SLO C11-5-14: Describe uses of aromatic hydrocarbons.
Examples: polychlorinated biphenyls, caffeine, steroids, organic solvents (toluene, xylene)...

Lesson Introduction

In the previous lesson, you learned the rules for naming and drawing alkynes. In this lesson, you will continue your study of the hydrocarbon family, this time with a focus on the aromatic hydrocarbons. You will examine the similarities and differences between the aromatic and aliphatic hydrocarbons, including some uses of aromatic hydrocarbons.

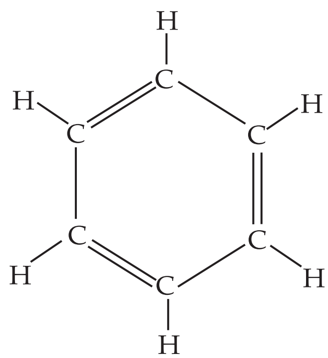
Defining Aromatic Hydrocarbons

There is a class of organic compounds that are responsible for the aromas of some common spices, like cinnamon, wintergreen, vanilla, cloves, and anise (liquorice). These compounds are called aromatic because some of those first discovered had pleasant odours. However, not ALL aromatic compounds have a scent. Aromatic compounds are also known as **arenes**.

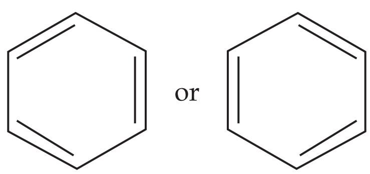
Molecules of aromatic compounds contain a single ring, or group of rings. Benzene, C_6H_6, is the simplest example of an aromatic compound. The benzene molecule consists of a six carbon ring with one hydrogen attached to each carbon atom. This leaves one electron from each carbon to participate in a double bond. The presence of three double bonds makes benzene both very stable and flat. If the benzene ring is a substituent of a hydrocarbon chain, it is called a **phenyl group**.

Discovery of Benzene

Benzene was first discovered by Michael Faraday in 1825. Its structure was somewhat of a mystery since it appeared to be an unsaturated compound but did not have the same chemical properties as other unsaturated hydrocarbons. In 1865, Friedrich August Kekulé dreamed about snakes moving around. One of the snakes grabbed its own tail and kept whirling before his eyes. From this, Kekulé proposed the structure of benzene to be a ring structure with alternating double bonds, like this:



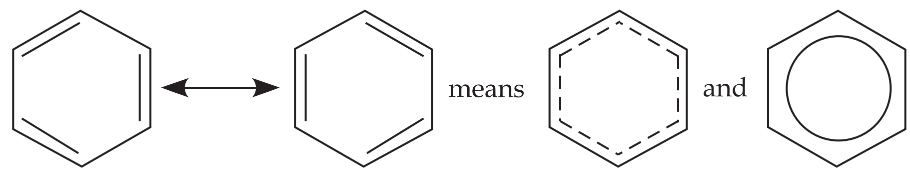
This structure can also be also drawn without showing the carbons and hydrogens, as follows:



Each vertex (corner) of the molecule is considered to be a carbon atom with a hydrogen. Both of the previous structures with alternating double bonds can be written for benzene. When two or more equally correct structures can be drawn for one molecule, it is said that **resonance** occurs. In the case of benzene, the actual bonds within the ring are hybrids of single and double bonds. This structure creates a very stable molecule, explaining why benzene is not as reactive as a six carbon alkene.

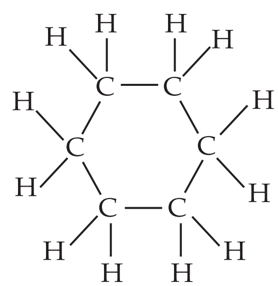
The structure suggested by Kekulé didn’t satisfy the chemical properties of benzene, although it was the best model available. It wasn't until 1945 that X-ray diffraction studies gave a clearer picture of benzene's structure. The studies indicated that only one structure for benzene existed, meaning that the resonance structures were not correct. Further studies showed that three pairs of electrons are actually shared equally by all six carbon atoms. It was determined that the carbon-carbon bonds are actually halfway between a single and double bond.

As the drawings above do not accurately represent the actual structure of benzene, the structure below was developed. The circle represents the three pairs of equally shared electrons.



Cyclic Hydrocarbons

Cyclic hydrocarbons are compounds that contain a hydrocarbon ring. They form when the two ends of a carbon chain are attached to form a ring. Five and six carbon rings are the most common. Any hydrocarbons that DO NOT possess rings are called aliphatic compounds. Here is an example of a cyclic hydrocarbon, cyclohexane:



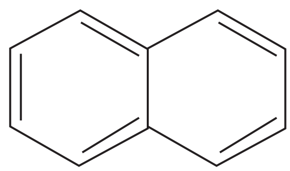
Note that cyclohexane has two fewer hydrogen atoms than hexane,

$CH_3CH_2CH_2CH_2CH_2CH_3$

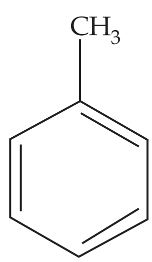
Why is this? This is due to the fact that when carbon atoms at the end of a chain bond to form a ring, those atoms can only form two additional bonds instead of three. You will note the same scenario occurs when other cyclic hydrocarbons form.

Uses of Aromatic Hydrocarbons

There are many examples of aromatic hydrocarbons that are used in everyday life. For example, *naphthalene* (also known as naphthaline, tar camphor, and antimite) is a crystalline, aromatic, white, solid hydrocarbon. Its formula is $C_{10}H_8$ and it has the structure of two fused benzene rings. Naphthalene is the primary ingredient of mothballs. It is volatile, forming a flammable vapour, and readily sublimes at room temperature, producing a characteristic odour that is detectable at very low concentrations. Here is the structure of naphthalene:



Another example of an aromatic hydrocarbon is *toluene*. Toluene, having the molecular formula C_7H_8, is a clear, colourless liquid at room temperature. It also has a sweet, pungent odour that is characteristic of benzene. Toluene is used extensively as a solvent and as a raw material in the manufacturing of several chemicals. It is also used in the rubber and plastics industries, and has even been used as a gasoline additive. Toluene is commonly used as a solvent or thinner in paints and lacquers. Here is the structure of toluene:



Learning Activity 6.7

Comparing Aromatic and Aliphatic Hydrocarbons

1. Cyclohexane and benzene both have six carbons bonded in a ring. What is the main difference between these two compounds?
2. What is a cyclic hydrocarbon?
3. Complete the following table:

	Aliphatic Hydrocarbon	**Aromatic Hydrocarbon**
Similarities		
Differences		

Lesson Summary

In this lesson, you learned that aromatic compounds contain the molecule benzene. The general molecular formula for benzene is C_6H_6. In the next lesson, you will learn to write structural formulas for and name common alcohols, including methyl, ethyl, and isopropyl alcohols.

NOTES

Assignment 6.7

Researching an Aromatic Compound *(12 marks)*

Choose one example of an aromatic compound. Do not use the two that were mentioned in this lesson (toluene and naphthalene). Examples of aromatic compounds that you can use include: polychlorinated biphenyls, caffeine, steroids, and organic solvents like xylene.

Once you have chosen an aromatic compound, you need to conduct some research in order to answer the eight questions below. Make sure that you keep track of the sources of your research, because you will be listing them in question 8.

1. Name of the aromatic compound. *(1 mark)*

   ____

   ____

   ____

2. List three uses of this aromatic compound. *(3 marks)*

   ____

   ____

   ____

3. Is this aromatic compound naturally occurring or synthetic? *(1 mark)*

   ____

   ____

   ____

continued

Assignment 6.7: Researching an Aromatic Compound (continued)

4. Draw the structure of the aromatic compound. *(1 mark)*

5. Write the molecular formula of the aromatic compound. *(1 mark)*

6. List the melting and boiling points of this compound. *(2 marks)*

7. Where is this compound found (if naturally occurring) or how is it made (if synthetic)? *(2 marks)*

8. List your sources of information. *(1 mark)*

continued

Assignment 6.7: Researching an Aromatic Compound (continued)

If you are not sure how to reference your sources of information, see the sample bibliography that follows:

Sample Bibliography

Book

Cochrane, Orin, et al. *Reading, Writing and Caring*. Winnipeg, MB: Whole Language Consultants, 1989.

Story, Poem, or Article from a Book

Foreword to the Future. Vol. 2. ed. Arthur Costa, James Bellanca, and Robin Fogarty. Sydney, Australia: Skylight Publishers, 1991. 52-64. (Example of an article)

Magazine Article

Lamb, Douglas H., and Glenn D. Reeder. "Reliving Golden Day." *Psychology Today* (June 1986): 22-25, 28.

Newspaper Article

Johnstone, Frederick. "People Cause Discrimination." *Winnipeg Free Press* 20 Dec. 1977: A3.

Website

Author (last name, first name). "Title of document." *Title of Complete Work* (if applicable). Date of publication to last date updated or revised (if known). <URL>. Date accessed—day, month, year.

Examples

Ignatius. "To the Trallians." *Early Church Documents* (*circa* 96-50 A.D.). 1994. http://listserv.american.edu/catholic/church/fathers/ignatius/ign-trl.txt. 20 Jun. 1996.

Anderson, Larry S. *Format for Update of Information Technology Plans.* 1994-1996. http://www2.msstate.edu/~lsal/nctp/index.html. 20 Jan. 1996.

NOTES

LESSON 8: COMMON ALCOHOLS (1 HOUR)

Lesson Focus

SLO C11-5-15: Write condensed structural formulas for and name common alcohols.
Include: maximum of six-carbon parent chain, IUPAC nomenclature

SLO C11-5-16: Describe uses of methyl, ethyl, and isopropyl alcohols.

Lesson Introduction

In the next several lessons, you are going to study substituted derivatives. These are organic compounds that replace one or more hydrogen atoms with another non-hydrocarbon group called a functional group. You are probably most familiar with the alcohols that are used in beverages, the gas tanks of cars, and in rubbing alcohol. In this lesson, you will study the structure and uses of several different alcohols.

The Hydroxyl Group

Functional groups are usually small units within an organic molecule that are responsible for most of the chemical and physical properties of that molecule. The functional group that identifies an alcohol is called the *hydroxyl group, –OH*.

The hydroxyl group should not be confused with the hydrox<u>ide</u> ion (OH^-) in an inorganic compound. The hydroxide ion is a charged particle that is joined to a positively charged ion. You would commonly see this when water dissociates in the following manner:

$$H_2O \rightarrow H^+ + OH^-$$

In another example, sodium hydroxide will dissociate into sodium ions and hydroxide ions:

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

The hydroxyl group in an organic compound is covalently bound to a carbon atom in the parent chain. In an alcohol, the hydroxyl does not dissociate from the rest of the organic compound, but rather the alcohol dissolves as a whole molecule:

$$R–OH_{(l)} \rightarrow R–OH_{(aq)}$$

The general molecular formula for an alcohol is *R–OH*, where the R represents a hydrocarbon chain.

Properties of Alcohols

The hydroxyl group is essentially one-half of a water molecule, so it will share some of the properties of water. Since the water molecule is a polar molecule, due to an unequal sharing of electrons in the covalent bonds, the hydroxyl group is also polar. The polarity of the hydroxyl group increases the boiling and melting points of alcohols, compared to similar hydrocarbons.

Compound	**Boiling Point (°C)**
methane, CH_4	–161.5
methanol, CH_3OH	65
ethane, C_2H_6	–88.6
ethanol, C_2H_5OH	78.5
propane, C_3H_8	–42.1
propan–1–o1, C_3H_7OH	97

The polar nature of the hydroxyl group allows smaller, polar alcohols to dissolve in polar water. Seeing as water and alcohol are both polar substances, like will dissolve like. *As the hydrocarbon chain in alcohols gets larger, the non-polar nature of the hydrocarbon chain begins to overcome the polar nature of the hydroxyl group, making larger alcohols insoluble in water. Generally, only alcohols with up to four carbons are soluble.*

Naming Alcohols

Many alcohols have common names, but you are going to name alcohols according to the IUPAC naming system. All alcohols end in *–ol*, which replaces the *–e* at the end of the name of the hydrocarbon parent chain. The longest carbon chain is numbered such that the carbon joined to the hydroxyl group is given the lowest number. That number is then infixed to the name of the alcohol.

Here again, the IUPAC naming system has been recently modified. You will likely notice that many textbooks and websites have not converted to the new naming rules. The old IUPAC names will be indicated for your reference, but on learning activities, assignments, and examinations you must apply the new naming rules.

Steps for Naming Alcohols

1. *Choose the longest continuous carbon chain that contains the hydroxyl group as the parent chain. Replace the –e at the end of the parent chain name with –ol.*
2. *Number the carbons so that the carbon attached to the hydroxyl group has the lowest number.*
3. *Infix the number of the carbon attached to the –OH group, separated with a hyphen.* Note that methanol and ethanol are unambiguous and do not require infixed position numbers.
4. *If there are any alkyl groups, name them as you did for the alkanes.*

Example 1

Give the IUPAC name for the following alcohol:

CH_3CH_2OH

Step 1: *Choose the longest continuous carbon chain that contains the hydroxyl group as the parent chain. Replace the –e at the end of the parent chain name with –ol.*

CH_3CH_2OH

The longest chain is two carbons, so this is *ethanol*.

Step 2: *Number the carbons so that the carbon containing the hydroxyl group has the lowest number.*

$\overset{1}{C}H_3\overset{2}{C}H_2OH$

Step 3: *Infix the number of the carbon attached to the –OH group, separated with a hyphen.*

Ethanol does not require an infixed numerical bonding position.

Step 4: *If there are any alkyl groups, name them as you did for the alkanes.*

There are no branches. Therefore, the name of this compound is *ethanol*.

Example 2

Give the IUPAC name for the following alcohol:

$$\begin{array}{l} \quad\;\; CH_3 \\ \quad\;\;\;\; | \\ CH_3CHCH_2CHCH_3 \\ \qquad\qquad\quad\; | \\ \qquad\qquad\;\; OH \end{array}$$

This compound could also be condensed to $CH_3CH(CH_3)CH_2CH(OH)CH_3$.

Step 1: *Choose the longest continuous carbon chain that contains the hydroxyl group as the parent chain. Replace the –e at the end of the parent chain name with –ol.*

$$\begin{array}{l} \quad\;\; CH_3 \\ \quad\;\;\;\; | \\ CH_3CHCH_2CHCH_3 \\ \qquad\qquad\quad\; | \\ \qquad\qquad\;\; OH \end{array}$$

The longest chain is five carbons, so this is *pentanol*.

Step 2: *Number the carbons so that the carbon containing the hydroxyl group has the lowest number.*

$$\begin{array}{l} \quad\;\; CH_3 \\ 5 \quad\; 4| \quad\; 3 \qquad 2 \quad 1 \\ CH_3CHCH_2CHCH_3 \\ \qquad\qquad\quad\; | \\ \qquad\qquad\;\; OH \end{array}$$

Step 3: *Infix the number of the carbon attached to the –OH group, separated with a hyphen.*

The hydroxyl is on the #2 carbon, so this is *pentan–2–ol*.

Step 4: *If there are any alkyl groups, name them as you did for the alkanes.*

There is a methyl group on the #4 carbon. Therefore, the name of this compound is *4–methylpentan–2–ol*. (According to the old IUPAC naming system, this would have been called *4–methyl–2–pentanol*.)

Common Alcohols

It was only in the 20th century that the IUPAC system of naming alcohols was developed. Alcohols had been known long before the naming system was developed, so the common names that had been given to each alcohol have endured.

Methanol (CH_3OH), the simplest alcohol, is commonly known as methyl alcohol. It is also known as wood alcohol because it can be produced from condensing the vapour from wood heated in the absence of air. This is called destructive distillation. When making homemade distilled alcohol, or moonshine, methanol is a common product. Drinking methanol can cause blindness, because it dissolves the layer covering the nerves from the eyes to the brain. Another common name for methanol is methyl hydrate, or gas line antifreeze, which is commonly added to gasoline tanks in the winter. Many high performance race cars use methanol as a fuel.

Ethanol (CH_3CH_2OH or C_2H_5OH) is also known as ethyl alcohol or grain alcohol. It is probably the most common alcohol, and it is formed from the fermentation of sugars. Ethanol dissolves in water as whole molecules:

$$C_2H_5OH_{(l)} \rightarrow C_2H_5OH_{(aq)}$$

Ethanol is the alcohol found in wine, beer, and other alcoholic beverages. It burns very cleanly, so it is added to gasoline to make the alternate fuel for your vehicle that is often called "gasohol."

Isopropyl or *rubbing alcohol* is used as a household disinfectant and to sterilize some medical instruments. The IUPAC name for rubbing alcohol is *propan–2–ol* (or *2–propanol* using the old IUPAC naming rules).

Ethylene glycol is car radiator antifreeze. It is a special type of alcohol that has two hydroxyl groups, called diols (the diol implies two hydroxyl groups). The IUPAC name for ethylene glycol is *ethan–1,2–diol* (or *1,2–ethandiol* according to the outdated IUPAC system). Its structure is shown below:

$$\begin{array}{ccc} CH_2 & - & CH_2 \\ | & & | \\ OH & & OH \end{array}$$

Glycerol, sometimes called glycerine, is an alcohol with three hydroxyl groups, called a triol. Glycerol is commonly found in hand soaps and lotion. The IUPAC name for glycerol is *propan–1,2,3–triol* (it used to be called *1,2,3–propanetriol*). The structure of glycerol is shown below. The three hydroxyl groups make glycerol very polar, and the strong forces of attraction between the molecules makes glycerol viscous, like honey. The hydroxyl groups tend to attract water as well, making it ideal as a moisturizer.

$$\begin{array}{ccccc} CH_2 & - & CH_2 & - & CH_2 \\ | & & | & & | \\ OH & & OH & & OH \end{array}$$

Learning Activity 6.8

Identifying Alcohols

1. Compare and contrast the effect of the OH group in organic and inorganic chemistry.
2. What are the IUPAC names for the following common alcohols?
   a) Rubbing alcohol
   b) Antifreeze
   c) Glycerine
3. What is the common ending when naming any alcohol? How is the parent chain numbered in relation to the hydroxyl group?
4. Identify which alcohol is most likely used to perform the following tasks:
   a) sterilizes medical equipment
   b) makes hand soap
   c) makes wine and beer
   d) makes gas line antifreeze
   e) used as an additive to gasoline for everyday cars

Lesson Summary

In this lesson, you learned that the functional group for alcohols is the hydroxyl group, *–OH*. The hydroxyl group is different from the hydroxide ion used in inorganic chemistry. You named alcohols by numbering the hydrocarbon parent chain (starting form the end closest to the hydroxyl group) and replaced the *–e* ending of the parent chain name with *–ol*. Next, you practiced drawing alcohols and discovered some uses of common alcohols. In the next lesson, you will write condensed structural formulas for and name organic acids.

NOTES

Assignment 6.8

Naming and Drawing Alcohols *(10 marks)*

1. What is the IUPAC name for each of the following alcohols? *(5 marks)*

a)
$$\begin{array}{l} OH \\ | \\ CH_2-CH_2-CH_3 \end{array}$$

____

b)
$$\begin{array}{lllll} OH & & & & \\ | & & & & \\ CH_2 & - & CH & - & CH_3 \\ & & | & & \\ & & CH_3 & & \end{array}$$

____

c)
$$\begin{array}{lllllllll} & & OH & & & & & & \\ & & | & & & & & & \\ CH_3 & - & C & - & CH_2 & - & CH & - & CH_3 \\ & & | & & & & | & & \\ & & CH_3 & & & & CH_3 & & \end{array}$$

____

d)
$$\begin{array}{lllllllllll} OH & & & & & & & & & & CH_3 \\ | & & & & & & & & & & | \\ CH_2 & - & CH_2 & - & CH_2 & - & CH_2 & - & CH_2 & - & CH_2 \end{array}$$

____

e) CH_3-CH_2-OH

or

$HO-CH_2-CH_3$

____

continued

Assignment 6.8: Naming and Drawing Alcohols (continued)

2. Draw the structural formula for each of the following alcohols. *(5 marks)*
   a) methanol

   b) hexan–3–ol

   c) 4–methylpentan–2–ol

   d) propan–1,2–diol

   e) 3,3–dimethylhexan–1–ol

Lesson 9: Organic Acids (1 hour)

Lesson Focus

SLO C11-5-17: Write condensed structural formulas for and name organic acids.
Include: maximum of six-carbon parent chain, IUPAC nomenclature

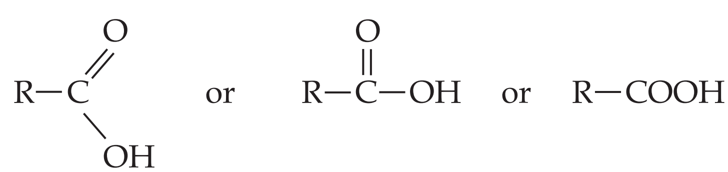
SLO C11-5-18: Describe uses or functions of common organic acids.
Examples: acetic, ascorbic, citric, formic, acetylsalicylic (ASA), lactic...

Lesson Introduction

Carboxylic acids are a group of organic acids responsible for the sour taste in citrus fruits and vinegar. A special kind of carboxylic acid, called an amino acid, joins together in complex chains to make up all proteins. Additionally, fats contain large carboxylic acids. In this lesson, we will examine the structure and properties of carboxylic acids.

The Carboxyl Group

The functional group of carboxylic acids is the *carboxyl group, –COOH*. The carboxyl group can be represented as:



Carboxylic acids contain one or more carboxyl groups. The general form for a carboxylic acid is R–COOH, where R represents a hydrocarbon chain.

Properties of Carboxylic Acids

The carboxyl group is very polar. *The polar nature of the carboxyl group overcomes the non-polar hydrocarbon parent chain, allowing smaller organic acids to be soluble in water. Like larger alcohols, the larger carboxylic acids are not soluble because the non-polar nature of the longer hydrocarbon parent chain overcomes the polar carboxyl group.*

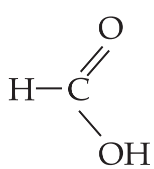
The polar nature of carboxylic acids gives them high melting and boiling points. For example, the boiling point of methanoic acid ($HCOOH$) is 101°C, and that of ethanoic acid (CH_3COOH) is 118°C.

Naming Carboxylic Acids

Like alcohols, many carboxylic acids have common names. When naming carboxylic acids with the IUPAC system, the parent hydrocarbon chain is named, replacing the *–e* ending with *–oic acid*. Numbering the parent chain always begins with the carbon containing the carboxyl group, and since the functional group is always attached to the first carbon, there is no need to write the "1."

Example 1

Name the following carboxylic acid:



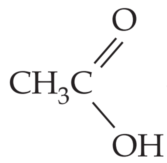
or $HCOOH$

Step 1: *Choose the longest continuous carbon chain that contains the carboxyl group as the parent chain. Replace the –e at the end of the parent chain name with –oic acid.*

There is a single carbon that would be called methane were it an alkane. The compound is then *methanoic acid*.

Example 2

Name the following carboxylic acid:



or CH_3COOH

Step 1: *Choose the longest continuous carbon chain that contains the carboxyl group as the parent chain. Replace the –e at the end of the parent chain name with –oic acid.*

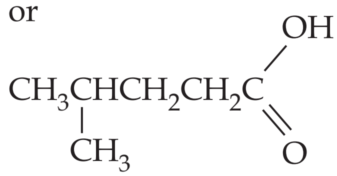
There are two carbons that would result in the compound being called ethane if this were an alkane. The compound is *ethanoic acid*.

Example 3

Name the following carboxylic acid:

$CH_3CH(CH_3)CH_2CH_2COOH$

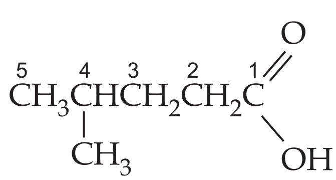
or



Step 1: *Choose the longest continuous carbon chain that contains the carboxyl group as the parent chain. Replace the –e at the end of the parent chain name with –oic acid.*

There are five carbons in the parent chain, so the compound would be called pentane if it were an alkane. The compound is *pentanoic acid*.

Step 2: *Number the carbons, beginning with the carboxyl group.*



Step 3: *If there are any alkyl groups, name them as you did for the alkanes.*

There is a methyl group attached to carbon #4. Therefore, the compound's name is *4–methylpentanoic acid*.

Common Carboxylic Acids

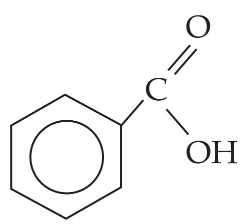
Methanoic acid, the simplest carboxylic acid, is commonly known as formic acid. It is called formic acid because it is responsible for the sting in ant bites (the Latin word for ant is *formica*).

Ethanoic acid is also known as *acetic acid*. You may recall that acetic acid is the acid that makes vinegar. The name acetic acid comes from the Latin word *acetum*, meaning sour.

Citric acid is an organic acid with three carboxyl groups found in citrus fruits like oranges and lemons. *Vitamin C*, or *ascorbic acid*, is also a carboxylic acid.

The IUPAC name for $CH_3CH_2CH_2COOH$ is *butanoic acid*. Its common name is *butyric acid*. Butanoic acid is responsible for the unpleasant smell of rancid butter, parmesan cheese, and vomit.

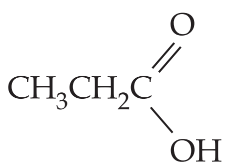
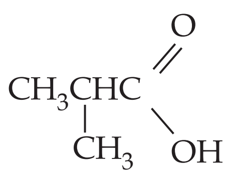
Benzoic acid, as its name suggests, contains a benzene ring. Benzoic acid is the substance that gives raspberries their tartness. It has the following structure:



Lactic acid, $CH_3CH(OH)COOH$, is formed when milk sours. Lactic acid is also the acid responsible for the burning sensation in your muscles when you exercise, due to glucose being converted into lactic acid.

Learning Activity 6.9

Naming Carboxylic Acids

1. Give the IUPAC name for each of the following acids.
   a) $CH_3CH_2CH_2CH_2CH_2COOH$
   b) 
   c) $CH_3C(CH_3)_2CH_2CH_2CH_2COOH$
   d) 
   e) $CH_3CH(C_2H_5)CH_2CH_2COOH$
2. Comment on the boiling points and melting points of carboxylic acids.
3. How does the polarity of the carboxyl group affect its solubility?
4. Identify which carboxylic acid is being described and give its common name.
   a) Causes the sting in ant bites
   b) Makes raspberries tart
   c) Causes the burn in muscles
   d) Makes vinegar sour
   e) Found in citrus fruits

Lesson Summary

In this lesson, you learned that carboxylic acids have the general form R—COOH, where R is a hydrocarbon chain. When naming carboxylic acids, the parent chain is named, with the carboxyl group being numbered first, dropping the *–e* ending of the parent chain name and adding *–oic acid*. In the next lesson, you will write condensed structural formulas, name, and describe the uses of some common esters.

Assignment 6.9

Drawing Carboxylic Acids *(5 marks)*

1. Draw the structure of each of the following acids. *(5 marks)*

   a) ethanoic acid

   ____

   b) 3–methylhexanoic acid

   ____

   c) 3,3–dimethylhexanoic acid

   ____

   d) 2,3–dimethylbutanoic acid

   ____

   e) 3–methylpentanoic acid

   ____

Notes

Lesson 10: Esters (1 hour)

Lesson Focus

SLO C11-5-20: Write condensed structural formulas for and name esters.
Include: up to 6-C alcohols and 6-C organic acids, IUPAC nomenclature

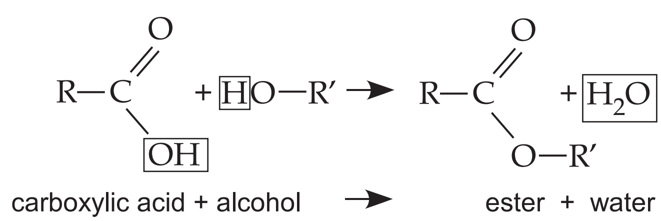
SLO C11-5-21: Describe uses or common esters.
Examples: pheromones, artificial flavourings...

Lesson Introduction

Many fruity flavours and scents are esters. In this lesson you will learn the rules for naming and writing the condensed formulas for esters. You will also discover some common esters that are used in your everyday life.

How are Esters Formed?

The process of forming an ester from the reaction between an organic acid and an alcohol is called **esterification**. In order for this reaction to occur, the acid and the alcohol must be heated in the presence of a catalyst, which is usually concentrated sulfuric acid.



In the reaction above, the –OH (hydroxyl group) from the acid combines with the H from the alcohol to form water. You will notice that the ester formed has two functional groups, R and R′. The molecular formula for an ester is *R–COO–R′* where the *R* represents the hydrocarbon chain from the acid and the *R′* represents the hydrocarbon chain from the alcohol.

Esters are polar molecules, but they do not form hydrogen bonds like water, alcohols, and carboxylic acids. *This means that the melting and boiling points of esters are higher than those of similar hydrocarbons, but lower than those of organic acids and alcohols. Their polar nature means esters are soluble in water.*

Naming Esters

Esters (R-COO-R′) are named as alkyl derivatives of carboxylic acids. When naming esters, the alkyl group from the alcohol (R′, it is joined to the singly bonded oxygen) is named first, replacing the *–ol* with *–yl*. In a separate word, the acid is named with the *–oic* acid ending replaced by *–oate*. Follow these steps to name esters:

Step 1: *Identify the groups from the acid and the alcohol.*

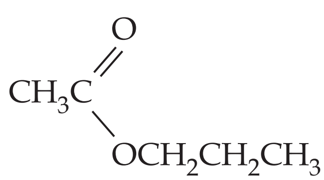
Step 2: *Name the alcohol (R′), replacing the –ol with –yl.*

Step 3: *Name the acid (R), replacing the –oic acid ending with –oate.*

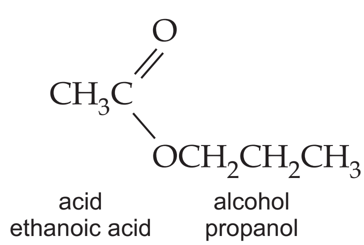
Step 4: *Put both parts of the name together.*

Example 1

Name the following ester:


or $CH_3COOCH_2CH_2CH_3$

Step 1: *Identify the groups from the acid and the alcohol.*



Step 2: *Name the alcohol (R′), replacing the –ol with –yl.*

The alcohol portion has three carbons, so its name is *propyl*.

Step 3: *Name the acid (R), replacing the –oic acid ending with –oate.*

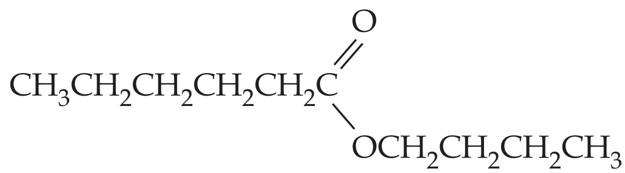
The acid portion has two carbons, so we name it *ethanoate*.

Step 4: *Put both parts of the name together.*

The name of the compound is *propyl ethanoate*.

Example 2

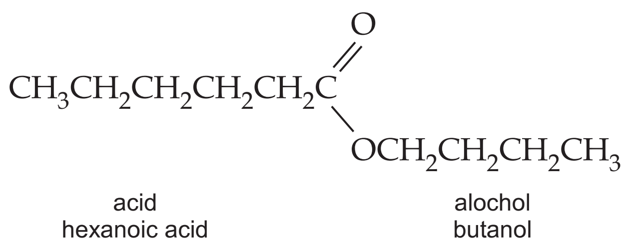
Give the IUPAC name for the following ester:



or

$CH_3CH_2CH_2CH_2CH_2COOCH_2CH_2CH_2CH_3$

Step 1: *Identify the groups from the acid and the alcohol.*



Step 2: *Name the alcohol (R′), replacing the –ol with –yl.*

The alcohol portion has four carbons, so its name is *butyl*.

Step 3: *Name the acid, replacing the –oic acid ending with –oate.*

The acid portion has six carbons, so its name is *hexanoate*.

Step 4: *Put both parts of the name together.*

The name of the compound is butyl *hexanoate*.

Example 3

Give the IUPAC name for the following ester:

$CH_3CH_2CH_2CH_2COOCH_3$

Step 1: *Identify the groups from the acid and the alcohol.*

There is methanol and pentanoic acid.

Step 2: *Name the alcohol (R′), replacing the –ol with –yl.*

Methanol becomes *methyl*.

Step 3: *Name the acid, replacing the –oic acid ending with –oate.*

Pentanoic acid becomes *pentanoate*.

Step 4: *Put both parts of the name together.*

This ester is called methyl *pentanoate*.

Drawing Esters

Esterification involves a carboxylic acid and an alcohol. When these are given as reactants, you can predict the products of the reaction and draw the resulting ester. Follow these steps to draw esters:

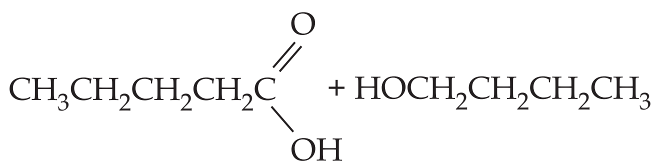
Step 1: *Draw the structures for the reactants (if they are not provided).*

Step 2: *Remove the -OH from the acid and the -H from the -OH of the alcohol, and join the two remaining structures. Then add water as the other product.*

Example 1

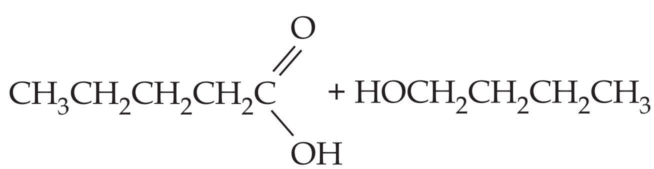
Write the esterification reaction for the reaction between pentanoic acid and butan-1-ol. Label and draw the structures for all reactants and products.

Step 1: *Draw the structures for the reactants.*

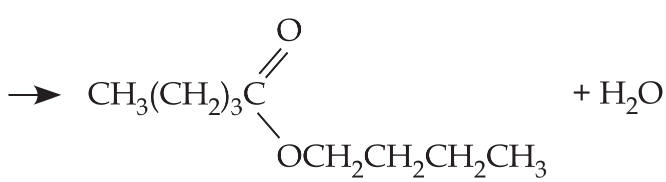


pentanoic acid + butan–1–ol

Step 2: *Remove the -OH from the acid and the -H from the -OH of the alcohol, and join the two structures. Then add water as the other product.*



pentanoic acid + butan–1–ol



→ butyl pentanoate + water

or

$CH_3CH_2CH_2CH_2COOH + HOCH_2CH_2CH_2CH_3$

$\rightarrow CH_3CH_2CH_2CH_2COOCH_2CH_2CH_2CH_3 + H_2O$

Esters: Flavours and Fragrances

Esters are a class of compounds widely distributed in nature. They have the following general formula:

$$\mathrm{R} - \overset{\overset{\displaystyle \mathrm{O}}{||}}{\mathrm{C}} - \mathrm{OR'}$$

The simpler esters tend to have pleasant odours. In many cases, although not exclusively so, the characteristic flavours and fragrances of flowers and fruits are due to compounds with the ester functional group. An exception is the case of the essential oils. The organoleptic qualities (odours and flavours) of fruits and flowers may often be due to a single ester, but, more often, the flavour or the aroma is due to a complex mixture in which a single ester predominates. Some common flavour principles are listed in Table 1. Food and beverage manufacturers are thoroughly familiar with these esters and often use them as additives to spruce up the flavour or odour of a dessert or beverage. Many times such flavours or odours are not even naturally occurring, as is the case with the "juicy fruit" principle, isopentenyl acetate. An instant pudding that has "rum" flavour may never have seen its alcoholic namesake—this flavour can be duplicated by the proper admixture, along with other minor components, of ethyl formate and isobutyl propionate. The natural flavour and odour are not exactly duplicated, but most people can be fooled. Often only a trained person with a high degree of gustatory perception—a professional taster—can tell the difference.

A single compound is rarely used in good quality imitation flavouring agents. A formula for an imitation pineapple flavour, which might fool an expert, is listed in Table 2. The formula includes ten esters and carboxylic acids, which may easily be synthesized in the laboratory. The remaining seven oils are isolated from natural sources.

Flavour is a combination of taste, sensation, and odour transmitted by receptors in the mouth (taste buds) and nose (olfactory receptors). The four basic tastes—sweet, sour, salty, and bitter—are perceived in specific areas of the tongue. The sides of the tongue perceive sour and salty tastes, the tip is most sensitive to sweet tastes, and the back of the tongue detects bitter tastes. The perception of flavour, however, is not that simple. If it were, it would only require the formulation of various combinations of four basic substances—a bitter substance (a base), a sour substance (an acid), a salty substance (sodium chloride), and a sweet substance (sugar)—to duplicate any flavour! In fact, we cannot duplicate flavours in this way. The human tongue actually possesses 9000 taste buds. It is a combined response of these taste buds that allows us to perceive a particular flavour.

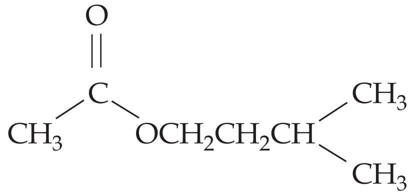
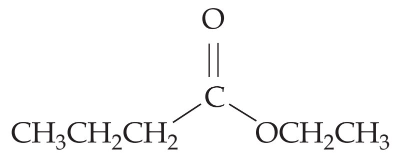
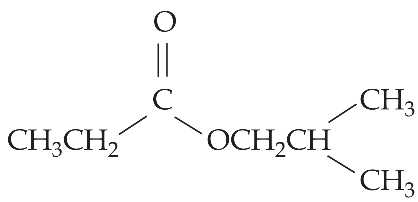
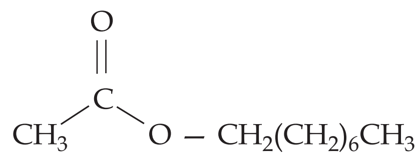
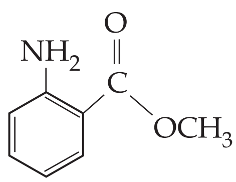
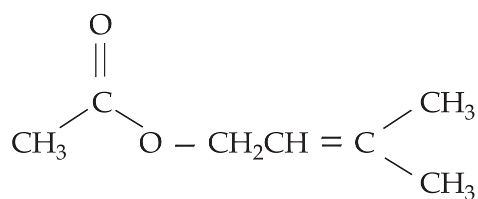
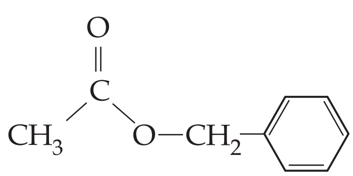
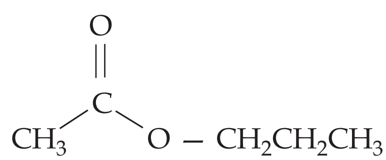
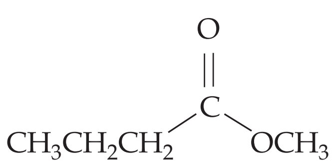
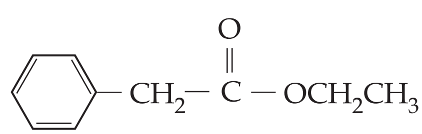
Table 1: Ester Flavours and Fragrances	
 **Isoamyl acetate** banana (alarm pheromone of honeybee)	 **Ethyl butyrate** pineapple
 **Isobutyl propionate** rum	 **Octyl acetate** oranges
 **Methyl anthranilate** grape	 **Isopentenyl acetate** "juicy fruit"
 **Benzyl acetate** peach	 **n-Propyl acetate** pear
 **Methyl butyrate** apple	 **Ethyl phenylacetate** honey

Table 2: Artificial Pineapple Flavour			
Pure Compounds	**%**	**Essential Oils**	**%**
Allyl caproate	5	Oil of sweet birch	1
Isoamyl acetate	3	Oil of spruce	2
Isoamyl isovalerate	3	Balsam Peru	4
Ethyl acetate	15	Volatile mustard oil	1
Ethyl butyrate	22	Oil cognac	5
Terpinyl propionate	3	Concentrated orange oil	4
Ethyl crotonate	5	Distilled oil of lime	2
Caproic acid	8		19
Butyric acid	12		
Acetic acid	5		
	81		

Although the "fruity" tastes and odours of esters are pleasant, they are seldom used in perfumes or scents that are applied to the body. The reason for this is a chemical one. The ester group is not as stable to perspiration as the ingredients of the more expensive essential oil perfumes. The latter are usually hydrocarbons (terpenes), ketones, and ethers extracted from natural sources. Esters, however, are only used for the cheapest toilet waters, since on contact with sweat, they undergo a hydrolysis reaction to give organic acids. These acids, unlike their precursor esters, generally do not have a pleasant odour. Butyric acid, for instance, has a strong odour similar to that of rancid butter (of which it is an ingredient) and is, in fact, a component of what we normally call "body odour." It is this substance that makes foul-smelling humans so easy for an animal to detect when it is downwind of them. It is also of great help to the bloodhound that is trained to follow small traces of this odour. The esters of butyric acid, ethyl butyrate and methyl butyrate, however, smell like pineapple and apple.

A sweet fruity odour also has the disadvantage that it may attract fruit flies and other insects in search of food. The case of isoamyl acetate, the familiar solvent called banana oil, is particularly interesting. It is identical to the alarm *pheromone* of the honeybee. Pheromone is the name applied to a chemical secreted by an organism that evokes a specific response in another member of the same species. This kind of communication is common between insects that otherwise lack means of intercourse. When a honeybee worker stings an intruder, an alarm pheromone, composed in part of isoamyl acetate, is secreted along with the sting venom. This chemical causes aggressive attack on the intruder by other bees, which swarm after the intruder. Obviously, it wouldn't be wise to wear a perfume compounded of isoamyl acetate near a beehive.

Learning Activity 6.10

Esterification Reactions

1. Complete the esterification equation for each of the following. Name and draw the structural formula for each reactant and product formed.

   a) $CH_3COOH + CH_3CH_2OH \rightarrow$

   b) $CH_3CH_2CH_2COOH + CH_3CH_2CH_2OH \rightarrow$

   c) $\_\_\_\_\_ + HOCH_2CH_3 \rightarrow CH_3CH_2CH_2\overset{O}{\overset{\|}{C}}OCH_2CH_3 + \_\_\_\_\_$

   d) $HCOOH + \_\_\_\_\_ \rightarrow H\overset{O}{\overset{\|}{C}}OCH_2CH_2CH_2CH_2CH_3 + \_\_\_\_\_$

   e) $\_\_\_\_\_ + \_\_\_\_\_ \rightarrow CH_3\overset{O}{\overset{\|}{C}}OCH_2CH_2CH_3 + \_\_\_\_\_$

   f) $\_\_\_\_\_ + \_\_\_\_\_ \rightarrow CH_3CH_2\overset{O}{\overset{\|}{C}}OCH_3 + \_\_\_\_\_$

2. Which esters are responsible for these flavours or fragrances?

   a) grape flavour

   b) “juicy fruit” flavour

   c) rum flavour

   d) The esters of butyric acid, a component of body odour, that smell like pineapple and apple

Lesson Summary

In this lesson, you learned that the general molecular formula for esters is RCOOR', where R is the parent chain of the acid and R' is the parent chain of the alcohol. You studied the rules for naming esters and practised using these rules. In the next lesson, you will describe the process of polymerization and identify important natural and synthetic polymers.

NOTES

Assignment 6.10

Naming and Drawing Esters *(16 marks)*

1. Name the following esters. *(4 marks)*

   a) $CH_3CH_2COOCH_2CH_2CH_2CH_3$

   ____

   b) $CH_3CH_2CH_2COOCH_2CH_2CH_2CH_2CH_2CH_3$

   ____

   c) $HCOOCH_2CH_3$

   ____

   d) $CH_3CH_2CH_2CH_2CH_2COOCH_2CH_2CH_3$

   ____

2. Complete the esterification equation for each of the following. Name and draw the structural formula for each reactant and product formed. *(12 marks)*

   a) $CH_3CH_2CH_2COOH + CH_3OH \rightarrow$

   b) $HCOOH + CH_3CH_2CH_2CH_2OH \rightarrow$

Notes

LESSON 11: CHEMISTRY AND INDUSTRY (1 HOUR)

Lesson Focus

SLO C11-5-22: Describe the process of polymerization and identify important natural and synthetic polymers.
Examples: polyethylene, polypropylene, polystyrene, polytetrafluoroethylene (Teflon®)

SLO C11-5-23: Describe how the products of organic chemistry have influenced quality of life.
Examples: synthetic rubber, nylon, medicines, polytetrafluoroethylene (Teflon®)

Lesson Introduction

Plastic bags, soft drink bottles, styrofoam cups, polyester, and bullet-proof vests are all examples of common synthetic materials called polymers. These are everyday items that make living our lives easier. In this lesson, you will learn how polymers are formed and how they have affected daily life.

What is a Polymer?

A **polymer** is a very large molecule made of many smaller repeating units called **monomers**. The word polymer comes from the Greek words "*poly*", meaning many, and "*mer*", meaning part.

Starch is an example of a naturally occurring polymer. Starch consists of many glucose molecules joined together in very long chains. You may have noticed that if you chew on a cracker long enough, it will begin to taste sweeter as the starch is broken down into simpler sugars. Other examples of naturally occurring polymers include cellulose (makes up plant cell walls and wood), silk, and proteins.

The first synthetic polymer was *nylon*, developed for the DuPont Corporation by Wallace Carothers in 1935. He mixed several organic substances together and found that they formed fibres. He perfected the process, and when nylon became available in 1939, it began replacing silk for making women's hosiery and parachutes.

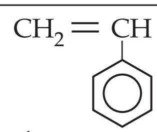
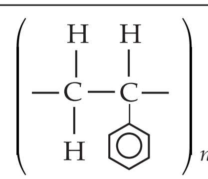
Addition Polymerization

There are two types of polymerization reactions—addition polymerization and condensation polymerization.

Addition polymers are formed by a reaction in which monomer units add together forming a long chain polymer. The monomer usually contains a carbon-carbon double bond. Polyethylene, polypropylene, Teflon™, Orlon, and synthetic rubbers are examples of polymeric products formed in this way.

Polymerization usually requires the presence of a small amount of **initiator**. Among the most common of these initiators are the peroxides. Chemists now know that the function of the peroxide is to produce a 'free radical.' It is these very reactive, unstable "free radicals" that add to alkenes to form larger alkane free radicals. This chain propagation forms the polymer. Eventually, two radicals combine and the process terminates.

Here are some common addition polymers:

Name of Polymer	Monomer	Repeating Structural Unit	Uses
polypropylene	$CH_2=CHCH_3$ propylene	$\left(-CH_2-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-\right)_n$	Moulded and extruded plastics, film, fibres for garments and carpeting
polystyrene	 styrene		Styrofoam, foam insulation, breakage-proof packaging, inexpensive moulded articles
polyvinyl chloride (PVC)	$CH_2=\underset{\displaystyle Cl}{\underset{\vert}{CH}}$ vinyl chloride	$\left(-\overset{\displaystyle H}{\overset{\vert}{\underset{\displaystyle H}{\underset{\vert}{C}}}}-\overset{\displaystyle H}{\overset{\vert}{\underset{\displaystyle Cl}{\underset{\vert}{C}}}}-\right)_n$	Electrical insulation, film, rubber-like articles, synthetic leather, floor covering, raincoats, shower curtains
polyfluoroethylene (Teflon™, PTFE)	$\overset{\displaystyle F}{\overset{\vert}{\underset{\displaystyle F}{\underset{\vert}{C}}}}=\overset{\displaystyle F}{\overset{\vert}{\underset{\displaystyle F}{\underset{\vert}{C}}}}$ tetrafluoroethene	$\left(-\overset{\displaystyle F}{\overset{\vert}{\underset{\displaystyle F}{\underset{\vert}{C}}}}-\overset{\displaystyle F}{\overset{\vert}{\underset{\displaystyle F}{\underset{\vert}{C}}}}-\right)_n$	Heat and stick resistant coating that is resistant to high temperatures and inert to almost all solvents and other chemicals

Condensation Polymerization

Condensation polymers are formed by the reaction of two or more different monomers. During the polymerization process, small molecules such as water, ammonia, or hydrogen chloride are produced. When developing nylon, Carothers noticed that water was produced and that the presence of water stopped the polymerization process. When he found a way to avoid water from entering the reaction vessel, he was able to form long nylon strands. Familiar examples of condensation polymers are nylon, Dacron, polyurethane and Silly Putty.

If you have access to the Internet, you can find out how rubber is like spaghetti by viewing a How Stuff Works video at http://dsc.discovery.com/videos/howstuffworks-rubber-and-spaghetti.html, how bubble gum can be made using a starch-polymer gum base at www.stevespanglerscience.com/experiment/00000133, how diapers take advantage of superabsorbent polymers at www.stevespanglerscience.com/experiment/00000064, and how scientists have begun to use polymers to clean up oil spills at www.stevespanglerscience.com/experiment/00000108. You can also learn more about polymers in general from the Macrogalleria website at www.pslc.ws/macrog/index.htm.

Try this—Making Your own Silly Putty™

Step 1: Pour 2 cups of white glue into a container (that is big enough to use your hands for mixing).

Step 2: Add food colouring to the glue.

Step 3: Mix 1 cup of starch to the glue using your hands. You can find starch at any grocery store.

Step 4: The mixing takes 10–15 minutes to reach the consistency of Silly Putty.

Step 5: Add more glue if the mixture seems too thick, or more starch if it seems too thin.

Step 6: Store the Silly Putty in an airtight container.

If you don't have the materials to make this Silly Putty, you can watch Steve Spangler make his own version at

www.stevespanglerscience.com/experiment/00000039.

Learning Activity 6.11

Comparing Polymers

Complete the following framework.

Type of Polymer	Addition	Condensation
List ways they are different from each other.		
Describe the process.		
Give examples of each type of polymer.		

Lesson Summary

In this lesson, you were introduced to the two types of polymerization. You also considered how the products of organic chemistry have influenced the quality of your life. In the next lesson, you will use the decision-making process to investigate an issue related to organic chemistry.

Assignment 6.11

Chemistry and Our Quality of Life *(10 marks)*

For this assignment, consider how the products of organic chemistry (e.g., synthetic rubber, nylon, medicines, and polytetrafluoroethylene [Teflon®] have influenced our quality of life. Then, complete the following rubric.

1. Are there only benefits to using these products? Support your answer with two supporting statements. *(2 marks)*

   (1 mark for stating position, 1 mark for each supporting reason)

2. Do these benefits outweigh any negative aspects associated with using these products? Explain. *(2 marks)*

   (1 mark for stating position, 1 mark for supporting statement)

continued

Assignment 6.11: Chemistry and Our Quality of Life (continued)

3. List three ways your life would be different if these products suddenly vanished. *(3 marks)*

   (1 mark for each point raised)

   ____

   ____

   ____

   ____

4. Should there be a penalty for not recycling? *(3 marks)*

   (1 mark for stating position, 1 mark for each supporting reason)

   ____

   ____

   ____

   ____

Lesson 12: Issues Related to Organic Chemistry (2 hours)

Lesson Focus

SLO C11-5-24: Use the decision-making process to investigate an issue related to organic chemistry.
Examples: gasohol production, alternative energy sources, recycling of plastics...

The Decision-Making Process

In order to complete this lesson, you may need to review the decision-making process. You may have used this model in other courses, such as Grade 10 Science. Essentially, the decision-making process involves obtaining the necessary background information on an issue, and then making an informed decision about the topic. To review this process, you can follow the example below, where the steps of this process are centred on a decision with which you are familiar:

Step 1: Identify the issue and raise a question.

Example:
Should you get your driver's license?

Step 2: Researching the implications of your decision.

What are the positive and negative effects of this decision?
Examples:
I may have more freedom. (positive)
Gas and parking will cost me money. (negative)

Step 3: Evaluate your research and develop possible courses of action.

Example:
You decide you do want your licence. Your choices are to go ahead and get your learner's permit or to first enrol in a local driving training course.

Step 4: Carefully make your decision and develop an action plan.

Example:
You decide to take a driver training course first.
You call several driving schools and book your session.

Step 5: Reflect on your decision and the process you used.

Example:
Did you consider all the options?
Are you happy with the decision?
What could you have done differently?

In this lesson, you will use the decision-making process to investigate an issue related to organic chemistry. Some examples of topics you may wish to explore include:

- The environmental impact of tar sands
- Alternative energy sources
- The recycling of plastics
- The production of methane from hog barns
- Industrial organic waste from local industry

Other topics are possible. Contact your tutor/marker if you have another idea that you think may be suitable.

While you have all the information from this course available to you, you will need additional research to completely investigate your issue. Other sources of information could include:

- Email communication with people in the community who have information on your topic
- Interviews
- Magazines
- Journals
- Newspapers
- Websites

Recording Information

A key element of this process is research so that you can take a position on an issue, and then ultimately make an informed decision. As such, it is important to keep all of your collected information organized. On the next page, you will find an example of a form that you can use to help organize and record information.

Lesson Summary

In this lesson, you used the following steps to arrive at a decision related to organic chemistry:

Step 1: *Identify an issue and raise a question.*

Step 2: *Research the implications of your decision.*

Step 3: *Evaluate your thoughts and develop possible courses of action.*

Step 4: *Carefully make your decision and develop an action plan.*

Step 5: *Reflect on your decision and the process you used to make your decision.*

As a result of this process, you presented your opinion on your chosen issue in the format of your choice.

This page is blank to accommodate the form on the following page.

Form for Recording Information	
Title of Source: ______ Author's Name: ______ Place of Publication: ______ Publisher: ______ Year of Publication: ______ Website Address: ______	
Summaries: Briefly note the main ideas of the text that you read.	**Paraphrases:** Write important and supporting information in your own words.
Questions: Write pertinent questions on what you have read or heard.	**Statistics:** Are there any recent statistics that influence your decision?
Comments: Record your own responses to questions about what you read.	**Direct Quotations:** Record quotes that are very important and are likely to be used in your supporting arguments for the decision you made.

Assignment 6.12

Investigating an Issue in Organic Chemistry *(16 marks)*

In this assignment, you will research and present your opinions on an issue of your choice, in the area of organic chemistry, using the decision-making model. Some examples of issues in organic chemistry include: gasohol production, alternative energy sources, and the recycling of plastics. Complete the following assignment to reflect on the process you used, and to elaborate upon your final decision. You should also look over the "Rubric for Assessment of Decision-Making Process Activity," so you know how you will be assessed on this assignment.

Step 1: Identify the issue you have chosen and raise a question.

____

____

____

____

____

Step 2: Research the implications of your decision. List as many points as you can.

____

____

____

____

____

continued

Assignment 6.12: Investigating an Issue in Organic Chemistry (continued)

Step 3: Suggest possible courses of action. List as many as you can.

Step 4: Carefully make your decision and develop an action plan.

My decision is:

My action plan is:

Step 5: Reflect on your decision and the process you used to make your decision. What could you have done differently?

Rubric for Assessment of Decision-Making Process Activity

Student Name(s) ______________________ Topic/Title ______________________

Criteria	Performance Levels			
	Level 1	**Level 2**	**Level 3**	**Level 4**
Selects an Option and Makes a Decision	❒ Is unable to come to a decision that clearly connects with the problem to be solved. ❒ Requires outside direction to make a choice.	❒ Identifies a feasible option, but cannot clearly decide on a plan. ❒ Requires outside influences to stand by a decision to proceed.	❒ Clearly selects an option and decides on a course of action, but others can identify that a better course of action remains untried. ❒ Recognizes potential safety concerns.	❒ Thoroughly analyzes all options collaboratively. ❒ Makes firm decision, justified by the research base, and recognizes most of the safety concerns.
Implements the Decision	❒ Is unable to implement the decision fully, but has an opportunity to modify it. ❒ Lacks the clarity to proceed.	❒ Implements the decision with a recognition that not all details are laid out in advance. ❒ Lacks clarity in having a plan for implementation.	❒ Implements the decision with some clarity of purpose. ❒ Demonstrates confidence that the implementation plan can follow a scientific inquiry approach.	❒ Implements the decision with clarity of purpose, backed by the research base. ❒ Clearly demonstrates that the implementation plan can be carried to completion as inquiry.
Identifies and Evaluates Actual Impacts of Decision	❒ Cannot clearly recognize more than one possible actual impact of the decision. ❒ Cannot effectively evaluate the effects of the decision taken.	❒ Can clearly recognize more than one possible actual impact of the decision taken. ❒ Cannot effectively evaluate the effects of the decision taken in most instances.	❒ Is able to recognize and comment upon the actual observed impacts of the decision. ❒ Demonstrates some ability to evaluate the impacts of the decision.	❒ Is able to recognize and comment deeply upon the actual observed impacts of the decision, noting unforeseen or unique outcomes. ❒ Is able to evaluate the impacts of the decision with ease.
Reflects on the Decision Making and Implementation of a Plan	❒ Begins to demonstrate an awareness of the need to review the implementation plan. ❒ Is reluctant to consider a re-evaluation of the plan.	❒ Reflects upon and intends to communicate the results of the implementation plan. ❒ Has some difficulty in knowing how to proceed with a re-evaluation of the problem-solving plan.	❒ Reflects upon and communicates the results of the implementation plan. ❒ Recognizes how to proceed with a re-evaluation of the problem-solving plan.	❒ Reaches higher order of synthesis in the reflection process. ❒ Has a sophisticated environmental awareness that informs this post-implementation period.

Note: *The above criteria are suggestions only, and will need to be adapted in collaboration with students according to the purpose of the assignment.*

NOTES

MODULE 6 SUMMARY

Congratulations! You have reached the end of Module 6. It is now time to write the final examination.

Submitting Your Assignments

It is now time for you to submit Assignments 6.1 to 6.12 to the Distance Learning Unit so that you can receive some feedback on how you are doing in this course. Remember that you must submit all the assignments in this course before you can receive your credit.

Make sure you have completed all parts of your Module 6 assignments and organize your material in the following order:

- Module 6 Cover Sheet (found at the end of the course Introduction)
- Assignment 6.1: Origins and Major Sources of Hydrocarbons
- Assignment 6.2: Aliphatic Hydrocarbons
- Assignment 6.3: Naming and Drawing Alkanes
- Assignment 6.4: Alkane Isomers*
- Assignment 6.5: Naming and Drawing Alkenes*
- Assignment 6.6: Naming and Drawing Alkynes*
- Assignment 6.7: Researching an Aromatic Compound
- Assignment 6.8: Naming and Drawing Alcohols
- Assignment 6.9: Drawing Carboxylic Acids
- Assignment 6.10: Naming and Drawing Esters
- Assignment 6.11: Chemistry and Our Quality of Life
- Assignment 6.12: Investigating an Issue in Organic Chemistry

* Remember to include the photographs you took in Assignments 6.4, 6.5, and 6.6. You can either mail or electronically submit these photographs. See the note on the following page.

For instructions on submitting your assignments, refer to How to Submit Assignments in the course Introduction.

Remember to include the photographs you took in Assignments 6.4, 6.5, and 6.6. You can either mail or electronically submit these photographs.

- If you took your pictures with a film camera or with a digital camera and printed them, you can mail them to the Distance Learning Unit along with the rest of the assignments.
- If you took your pictures with a digital camera, you can electronically submit them, as long as the files are smaller than 5 MB. If the files are larger than 5 MB, you must submit them as compressed files.

Final Examination

Congratulations, you have finished Module 6 in the course. The final examination is out of 100 marks and worth 30% of your final mark. In order to do well on this examination, you should review all of your learning activities and assignments from Modules 1 to 6. Please note that 80–85 percent of the final examination is concentrated on Modules 4 to 6.

You will complete this examination while being supervised by a proctor. You should already have made arrangements to have the examination sent to the proctor from the Distance Learning Unit. If you have not yet made arrangements to write it, then do so now. The instructions for doing so are provided in the Introduction to this module.

You will need to bring the following items to the examination: some pens and/or pencils (2 or 3 of each), an eraser or correction fluid, some blank paper, a ruler, and a scientific or graphing calculator.

A maximum of 2.5 hours is available to complete your final examination. When you have completed it, the proctor will then forward it for assessment. Good luck!

Grade 11 Chemistry (30S)

Module 6: Organic Chemistry

Learning Activity Answer Keys

Module 6: Organic Chemistry

Learning Activity 6.1: Organic Chemistry Introduction

1. Explain the two ways that chemists classify substances.

   Answer:

   Substances are classified as being either:

   Organic—made from living matter.

   Inorganic—not made from living matter.

2. How did Friedrich Wöhler change the way scientists perceived the nature of organic compounds?

   Answer:

   Up until the early 1800s, scientists believed that the ability to produce organic compounds rested exclusively with living things. Friedrich Wöhler changed this thinking in 1828 when he created urea in the absence of living things.

3. Define the term "organic compound," according to the findings of Friedrich Wöhler.

   Answer:

   The definition of organic compounds has since been extended to carbon-containing compounds, regardless of how they are created.

4. Name and briefly describe the two steps involved in the refining of crude oil.

   Answer:

   Step 1—Distillation: The crude oil is heated, creating vapours that move to a fractionating tower. When the vapours reach the temperature at which they condense, the liquid is collected at various points up the tower. Each "fraction" removed is further purified and refined.

   Step 2—Cracking: Cracking is a process where larger molecules are broken into smaller, more useable molecules. The products of cracking are used in the making of gasoline and starting molecules for synthetic petroleum products.

5. List three major sources of hydrocarbons.

   Answer:

   Most hydrocarbons are extracted from petroleum.

   Naturally occurring organic compounds result from the decay of prehistoric animals and vegetation.

   Other major sources of petroleum products are those that are synthetically produced.

6. List three distinguishing properties of hydrocarbons.

   Answer:

   All hydrocarbons are non-polar molecules.

   Hydrocarbons are insoluble in water.

   Hydrocarbons molecules are held together by weak intermolecular forces.

   Hydrocarbons have very low melting and boiling points, in relation to their mass.

   Hydrocarbons will burn in oxygen to form carbon dioxide and water.

7. What is petroleum?

   Answer:

   Petroleum is a mixture of over a thousand different organic compounds. These compounds can be separated based upon their boiling points using a process called fractional distillation.

Learning Activity 6.2: Properties of Normal Alkanes

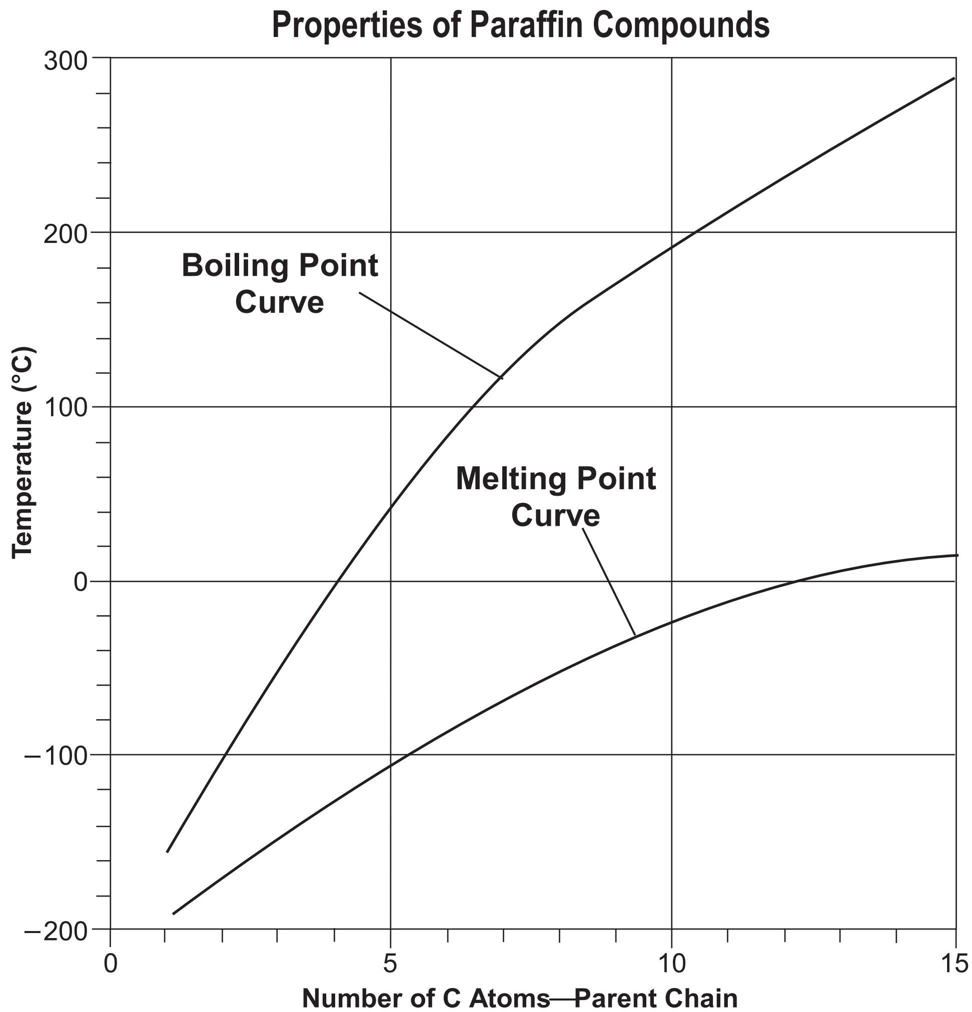
A paraffin compound is an example of an alkane. Remember that alkanes are saturated hydrocarbons having the general formula C_nH_{2n+2}, (where C is a carbon atom, H is a hydrogen atom, and n is the number of carbons). The paraffins are major constituents of natural gas and petroleum. Paraffins containing fewer than five carbon atoms per molecule are usually gaseous at room temperature, while those having five to 15 carbon atoms are usually liquids, and the straight chain paraffins having more than fifteen carbon atoms per molecule are solids.

In this learning activity, you will plot a graph to illustrate the relationship between the number of carbons in the parent chain of an alkane and its melting and boiling point.

Properties of Paraffin Hydrocarbons		
Number of Carbons in Parent Chain	**Melting Point (°C)**	**Boiling Point (°C)**
1	–180	–160
2	–190	–90
3	–175	–40
4	–140	0
5	–140	40
6	–90	80
7	–95	100
8	–50	140
9	–45	150
10	–20	180
11	–20	195
12	0	220
13	–5	250
14	5	270
15	15	290

1. Use the data table provided to complete the graph below. Plot *temperature* on the *y*-axis and *number of carbon atoms* on the *x*-axis. Your graph will show two curves, one for boiling point and one for melting point. To differentiate the two curves, use two different colours or draw one curve using a solid line and the other curve using a broken line.

   Answer:



2. What relationship exists among melting point (MP), boiling point (BP), and the number of carbons in the parent chain?

   Answer:

   The MP and BP are proportional to the length of the hydrocarbon chain. For example, there is a noticeable increase in the values of MP and BP for carbon chains from 1-C to 8-C.

Learning Activity 6.3: Working with Alkanes

1. How long can an alkyl group be? Draw and name the first three alkyl groups.

   Answer:

   An alkyl group can be one to several carbons long.

   The first three are the methyl group (CH_3-), the ethyl group (CH_3CH_2-), and the propyl group ($CH_3CH_2CH_2-$).

2. Name and draw the simplest straight chain alkane.

   Answer:

   Ethane

   CH_3CH_3

3. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it. (**Hint:** Draw the structures)

   a) 5–methylhexane

      Answer:

      Alkyl groups must be identified by the lowest number. The correct name is 2–methylhexane.

   b) 2–ethyl–2–methylpropane

      Answer:

      You must find the longest continuous carbon chain in the compound. The correct name is 2,2–dimethylbutane.

   c) 1–methylbutane

      Answer:

      You must find the longest continuous carbon chain in the compound. The correct name is pentane.

   d) 3–ethylbutane

      Answer:

      You must find the longest continuous carbon chain in the compound. The correct name is 3–methylpentane.

   e) 4–methyloctane

      Answer:

      The name is correct.

f) 2,3–trimethylheptane

Answer:

There are only two alkyl groups, so the name can't start with "trimethyl." The correct name is 2,3–dimethylheptane.

g) 1,4–dimethylpentane

Answer:

You must find the longest continuous carbon chain in the compound. The correct name is 2–methylhexane.

h) 1,3,5–trimethylpentane

Answer:

You cannot have a methyl group on a #1 carbon. The correct name is 4-methylheptane.

4. Name each of the following:

a) $CH_3CH_2CH_3$

Answer:

propane

b) $CH_3CH(CH_3)CH_2CH_2CH_3$

Answer:

2–methylpentane

c) $CH_3CH(CH_3)CH_2CH(CH_2CH_3)CH_2CH_2CH_3$

Answer:

4–ethyl–2–methylheptane

d)

$$\begin{array}{c} CH_3 \\ | \\ CH_3CCH_3 \\ | \\ CH_3 \end{array}$$

Answer:

2,2–dimethylpropane

e) $$\begin{array}{ccccc} & CH_3 & & CH_3 & \\ & | & & | & \\ CH_3 & C & CH_2 & CH & CH_3 \\ & | & & & \\ & CH_3 & & & \end{array}$$

Answer:

2,2,4–trimethylpentane

f) $CH_3CH_2C(CH_3)_2CH_2CH_3$

Answer:

3,3–dimethylpentane

5. Draw the structural formulas for each of the following, condensing when possible:

a) butane

Answer:

$CH_3CH_2CH_2CH_3$

b) octane

Answer:

$CH_3(CH_2)_6CH_3$

c) 2,3-dimethylpentane

Answer:

$$\begin{array}{ccccc} & CH_3 & & & \\ & | & & & \\ CH_3 & CH & CH & CH_2 & CH_3 \\ & & | & & \\ & & CH_3 & & \end{array}$$

d) 3-ethylpentane

Answer:

$$\begin{array}{ccccc} & & CH_2CH_3 & & \\ & & | & & \\ CH_3 & CH_2 & CH & CH_2 & CH_3 \end{array}$$

e) 3-methylhexane

Answer:

$$\begin{array}{cccccc} & & CH_3 & & & \\ & & | & & & \\ CH_3 & CH_2 & CH & CH_2 & CH_2 & CH_3 \end{array}$$

f) 2,3–dimethylbutane

Answer:

$$
\begin{array}{cccc}
 & CH_3 & & \\
 & | & & \\
CH_3 & CH & CH & CH_3 \\
 & & | & \\
 & & CH_3 &
\end{array}
$$

g) 3–ethyl–2–methylpentane

Answer:

$$
\begin{array}{ccccc}
 & & CH_2CH_3 & & \\
 & & | & & \\
CH_3 & CH & CH & CH_2 & CH_3 \\
 & | & & & \\
 & CH_3 & & &
\end{array}
$$

h) 3,3–diethylheptane

Answer:

$$
\begin{array}{ccccccc}
 & & CH_2CH_3 & & & & \\
 & & | & & & & \\
CH_3 & CH_2 & C & CH_2 & CH_2 & CH_2 & CH_3 \\
 & & | & & & & \\
 & & CH_2CH_3 & & & &
\end{array}
$$

i) 2,3,5–trimethylhexane

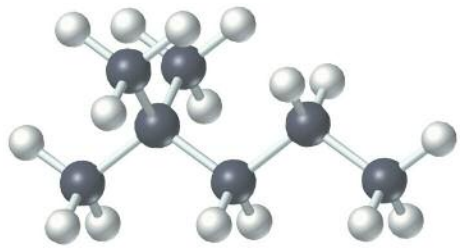
Answer:

$$
\begin{array}{cccccc}
 & CH_3 & & & & \\
 & | & & & & \\
CH_3 & CH & CH & CH_2 & CH & CH_3 \\
 & & | & & | & \\
 & & CH_3 & & CH_3 &
\end{array}
$$

6. Construct 3-D models for the following alkanes:

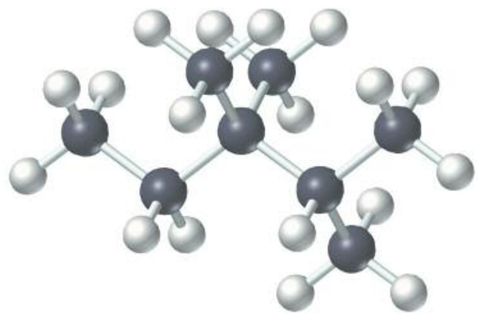
a) 2,2–dimethlypentane

Answer:



b) 2,3,3–trimethylpentane

Answer:



c) methane

Answer:

Learning Activity 6.4: Identifying Isomers of Alkanes

1. Indicate which straight chain alkane is a structural isomer of each of the following:

   a) 3–propylheptane

      Answer:

      10 carbons = decane

   b) 3–methylpentane

      Answer:

      6 carbons = hexane

   c) 2,2,3,3–tetramethylpentane

      Answer:

      9 carbons = nonane

   d) 2–methylbutane

      Answer:

      5 carbons = pentane

   e) 3–ethylhexane

      Answer:

      8 carbons = octane

   f) 2–methylhexane

      Answer:

      7 carbons = heptane

   g) 2,2–dimethylpentane

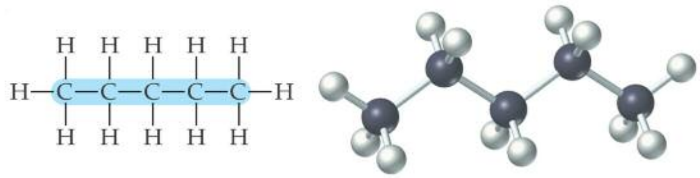
      Answer:

      7 carbons = heptane

2. Define "structural isomer." Next, name and draw two compounds that are structural isomers.

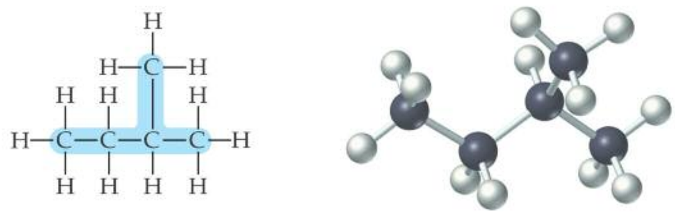
   Answer:

   Structural isomers are molecules that have the same molecular formula but different structural formulas. Examples will vary, but may include:



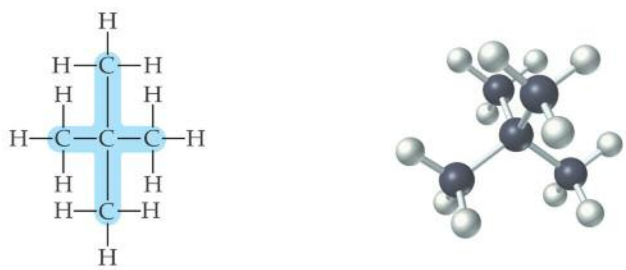
Straight chain

pentane



Branched chain

2–methylbutane



Branched chain

2,2–dimethylpropane

Learning Activity 6.5: Identifying Alkenes

1. What is the difference between saturated and unsaturated organic compounds? Give an example of each.

   Answer:

   Saturated and unsaturated refers to the possible number of hydrogen atoms in an organic compound. An organic compound that contains the maximum number of hydrogen atoms is saturated. An alkane is an example. If the compound does not contain as many hydrogen atoms as possible, it is unsaturated and contains at least one double or triple bond. An alkene is an example.

2. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it.

   a) 2–methylbut–3–ene

   Answer:

   $$\begin{array}{l} CH_3CHCH{=}CH_2 \\ \quad\;\;\, | \\ \quad\;\; CH_3 \end{array}$$

   The name of the compound should be 3–methylbut–1–ene, since the double bond must be given the lowest number.

   b) 2,2–dimethylpent–3–ene

   Answer:

   $$\begin{array}{l} \quad\;\; CH_3 \\ \quad\;\;\, | \\ CH_3CCH{=}CHCH_3 \\ \quad\;\;\, | \\ \quad\;\; CH_3 \end{array}$$

   The name of the compound should be 4,4–dimethylpent–2–ene.

   c) 1–butene

   Answer:

   The compound’s correct name is but–1–ene.

d) 2,2–dimethylhex–1–ene

Answer:

$$\begin{array}{l} CH_3 \\ \,| \\ H_2C=C-CH_2CH_2CH_2CH_3 \\ \,| \\ CH_3 \end{array}$$

This structure is not possible since the #2 carbon has five bonds.

e) 3–methylprop–1–ene

Answer:

This compound would actually be named but–1–ene, as it creates a parent chain of four carbons.

f) hex–5–ene

Answer:

$CH_3CH_2CH_2CH_2CH=CH_2$

The compound should be called hex-1-ene, since the double bond should get the lowest number.

3. Write the balanced equation, using structural formulas, for the following reactions:

a) The hydrogenation of but–2–ene

Answer:

$CH_3CH=CHCH_3 + H_2 \rightarrow CH_3CH_2CH_2CH_3$

b) The dehydrogenation of propane

Answer:

$CH_3CH_2CH_3 \rightarrow CH_3CH= CH_2 + H_2$

c) 4–methylhex–2–ene + $1H_2$

Answer:

$CH_3CH=CHCH(CH_3)CH_2CH_3 + H_2 \rightarrow CH_3CH_2CH_2CH(CH_3)CH_2CH_3$

4. Name the products of the chemical equations in question #3 above.

Answer:

a) butane

b) prop–1–ene

c) 3–methylhexane

Learning Activity 6.6: Identifying Alkynes

1. What, if anything, is wrong with the following names? If the compound exists, write the correct name for it.

   a) hex–4–yne

      Answer:

      The triple bond should get the lowest number, so the compound should be named hex–2–yne.

   b) 2,2–dimethyl–3–pentyne

      Answer:

      The triple bond should get the lowest number, so the compound should be named 4,4–dimethylpent–2–yne.

   c) but–1–yne

      Answer:

      The compound's name is correct.

   d) 2,2–dimethylhex–2–yne

      Answer:

      This structure is not possible since there are six bonds on the #2 carbon.

   e) 3–methylprop–2–yne

      Answer:

      The structure should be called but–2–yne, since you must find the longest continuous carbon chain in the compound that contains the triple bond.

2. Write the balanced equation, using structural formulas, for the following reactions:

   a) hydrogenation of but–2–yne

      Answer:

      $CH_3C\equiv CCH_3 + H_2 \rightarrow CH_3CH=CHCH_3$

   b) dehydrogenation of prop–1–ene

      Answer:

      $CH_2=CHCH_3 \rightarrow CH\equiv CCH_2 + H_2$

   c) 4–methylhex–2–yne + $1H_2$

      Answer:

      $CH_3C\equiv CCH(CH_3)CH_2CH_3 + H_2 \rightarrow CH_3CH=CHCH(CH_3)$

3. Name the products formed in question 2 above.

   Answer:

   a) but–2–ene
   b) prop–1–yne
   c) 4–methylhex–2–ene

Learning Activity 6.7: Comparing Aromatic and Aliphatic Hydrocarbons

1. Cyclohexane and benzene both have six carbons bonded in a ring. What is the main difference between these two compounds?

   Answer:

   Cyclohexane is a saturated cyclic hydrocarbon (meaning a carbon chain in the form of a ring that contains the maximum possible number of hydrogens), while benzene is an unsaturated cyclic hydrocarbon (called an aromatic hydrocarbon).

2. What is a cyclic hydrocarbon?

   Answer:

   A carbon chain in the form of a ring.

3. Complete the following table:

   Answer:

	Aliphatic Hydrocarbon	**Aromatic Hydrocarbon**
Similarities	Can be saturated or unsaturated.	Unsaturated.
Differences	Does not contain a benzene ring, does not show resonance.	Contains a benzene ring, demonstrates resonance.

Learning Activity 6.8: Identifying Alcohols

1. Compare and contrast the effect of the OH group in organic and inorganic chemistry.

   Answer:

Same (Compare)	Different (Contrast)
Both are represented by OH. Both are polar.	In ionic compounds, the hydroxide is an ion. In organic compounds, the hydroxyl group is neutral. In ionic compounds, the hydroxide ion is joined to a metal ion, while in organic compounds, the hydroxyl group is covalently joined to a carbon. In ionic compounds, the hydroxide ion dissociates in water, while in organic compounds, the hydroxyl group remains a molecule.

2. What are the IUPAC names for the following common alcohols?

   a) Rubbing alcohol

      Answer:

      propan–2–ol

   b) Antifreeze

      Answer:

      ethan–1,2–diol

   c) Glycerine

      Answer:

      propan–1,2,3–triol

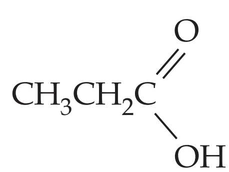
3. What is the common ending when naming any alcohol? How is the parent chain numbered in relation to the hydroxyl group?

   Answer:

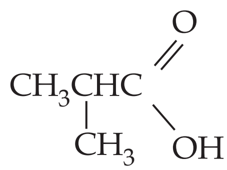
   All alcohols end in *–ol*, which replaces the *–e* at the end of the name of the parent hydrocarbon chain that contains the hydroxyl group. The parent chain is numbered such that the carbon joined to the hydroxyl group is given the lowest number. That number is then infixed into the name of the alcohol.

4. Identify which alcohol is most likely used to perform the following tasks:
   a) sterilizes medical equipment
      Answer:
      isopropyl / propan–2–ol
   b) makes hand soap
      Answer:
      glycerol / propan–1,2,3–triol
   c) makes wine and beer
      Answer:
      ethanol
   d) makes gas line antifreeze
      Answer:
      methanol
   e) used as an additive to gasoline for everyday cars
      Answer:
      ethanol

Learning Activity 6.9: Naming Carboxylic Acids

1. Give the IUPAC name for each of the following acids.
   a) $CH_3CH_2CH_2CH_2CH_2COOH$
      Answer:
      hexanoic acid
   b) 

      Answer:
      propanoic acid
   c) $CH_3C(CH_3)_2CH_2CH_2CH_2COOH$
      Answer:
      5,5–dimethylhexanoic acid

d) 

Answer:

2–methylpropanoic acid

e) $CH_3CH(C_2H_5)CH_2CH_2COOH$

Answer:

4–ethylhexanoic acid

2. Comment on the boiling points and melting points of carboxylic acids.

   Answer:

   The polar nature of the carboxylic acids gives them high melting and boiling points. For example, the boiling point of methanoic acid (HCOOH) is 101°C and that of ethanoic acid (CH_3COOH) is 118°C.

3. How does the polarity of the carboxyl group affect its solubility?

   Answer:

   The carboxyl group is very polar. The polar nature of the carboxyl group overcomes the non-polar hydrocarbon parent chain, allowing smaller organic acids to be soluble in water. Like larger alcohols, the larger carboxylic acids are not soluble because the non-polar nature of the longer hydrocarbon parent chain overcomes the polar carboxyl group.

4. Identify which carboxylic acid is being described and give its common name.

   a) Causes the sting in ant bites

   Answer:

   formic acid

   b) Makes raspberries tart

   Answer:

   benzoic acid

   c) Causes the burn in muscles

   Answer:

   lactic acid

d) Makes vinegar sour

Answer:

acetic acid

e)) Found in citrus fruits

Answer:

citric acid

Learning Activity 6.10: Esterification Reactions

1. Complete the esterification equation for each of the following. Name and draw the structural formula for each reactant and product formed.

a) $CH_3COOH + CH_3CH_2OH \rightarrow$

Answer:

$CH_3COOH + CH_3CH_2OH \rightarrow CH_3COOCH_2CH_3 + H_2O$
ethanoic acid ethanol ethyl ethanoate water

b) $CH_3CH_2CH_2COOH + CH_3CH_2CH_2OH \rightarrow$

Answer:

$CH_3CH_2CH_2COOH + CH_3CH_2CH_2OH \rightarrow$
butanoic acid propan–1–ol

$CH_3CH_2CH_2COOCH_2CH_2CH_3 + H_2O$
propyl butanoate water

c)

$$\_\_\_\_\_ + HOCH_2CH_3 \rightarrow CH_3CH_2CH_2\overset{\overset{O}{||}}{C}OCH_2CH_3 + \_\_\_\_\_$$

Answer:

$$CH_3CH_2CH_2\overset{\overset{O}{||}}{C}-OH + HOCH_2CH_3 \rightarrow CH_3CH_2CH_2\overset{\overset{O}{||}}{C}OCH_2CH_3 + H_2O$$

butanoic acid ethanol ethyl
butanoate water

d)

$$HCOOH + \_\_\_\_\_ \rightarrow H\overset{\overset{O}{||}}{C}OCH_2CH_2CH_2CH_2CH_3 + \_\_\_\_\_$$

Answer:

$$H\overset{\overset{O}{||}}{C}-OH + HOCH_2CH_2CH_2CH_2CH_3 \rightarrow H\overset{\overset{O}{||}}{C}OCH_2CH_2CH_2CH_2CH_3 + H_2O$$

methanoic acid pentan–1–ol pentyl methanoate water e)

$$\_\_\_\_\_ + \_\_\_\_\_ \rightarrow CH_3\overset{\overset{\displaystyle O}{||}}{C}OCH_2CH_2CH_3 + \_\_\_\_\_$$

Answer:

$$\underset{\text{ethanoic acid}}{CH_3\overset{\overset{\displaystyle O}{||}}{C}-OH} + \underset{\text{propan–1–ol}}{HOCH_2CH_2CH_3} \rightarrow \underset{\text{propyl methanoate}}{CH_3\overset{\overset{\displaystyle O}{||}}{C}OCH_2CH_2CH_3} + \underset{\text{water}}{H_2O}$$

f)

$$\_\_\_\_\_ + \_\_\_\_\_ \rightarrow CH_3CH_2\overset{\overset{\displaystyle O}{||}}{C}OCH_3 + \_\_\_\_\_$$

Answer:

$$\underset{\text{propanoic acid}}{CH_3CH_2\overset{\overset{\displaystyle O}{||}}{C}-OH} + \underset{\text{methanol}}{CH_3OH} \rightarrow \underset{\text{methyl propanoate}}{CH_3CH_2\overset{\overset{\displaystyle O}{||}}{C}OCH_3} + \underset{\text{water}}{H_2O}$$

2. Which esters are responsible for these flavours or fragrances?

a) grape flavour

Answer:

methyl anthranilate

b) "juicy fruit" flavour

Answer:

isopentenyl acetate

c) rum flavour

Answer:

ethyl formate and isobutyl propionate

d) The esters of butyric acid, a component of body odour, that smell like pineapple and apple

Answer:

ethyl butyrate and methyl butyrate

Learning Activity 6.11: Comparing Polymers

Complete the following framework.

Answer:

Type of Polymer	Addition	Condensation
List ways they are different from each other.	Addition polymers start with a single type of monomer. Addition polymerization requires a free radical initiator.	Condensation polymers begin with more than one kind of monomer. Condensation polymerization does not require a free radical initiator.
Describe the process.	A monomer unit adds together forming a long chain polymer. The monomer usually contains a carbon-carbon double bond. The result is usually a long chain alkane. Addition polymerization does not generally produce any other particles.	Condensation polymers are formed by the reaction of two or more different monomers. During the condensation polymerization process, small molecules such as water, ammonia, or hydrogen chloride are produced.
Give examples of each type of polymer.	Polyethylene Polypropylene Polystyrene Polyvinyl chloride	Nylon Dacron (polyester)

NOTES

Grade 11 Chemistry (30S)

Appendices

Appendix A: Glossary

Absolute zero: –273.15°C. Theoretically the lowest possible temperature, where all molecular motion would cease, the kinetic energy would be zero, and the volume of a gas, hypothetically, would also be zero.

Actual yield: The amount of product that is actually formed during a chemical reaction.

Addition polymers: Formed by a reaction in which monomer units add together forming a long polymer chain. The monomer usually contains a carbon-carbon double bond.

Aliphatic hydrocarbons: Organic compounds that have carbon atoms arranged in straight or branched chains, they can be divided into alkanes, alkenes, and alkynes.

Alkanes: Organic compounds containing only single carbon-carbon bonds, where the *carbon* atoms share only one pair of electrons. The general molecular formula for alkanes is C_nH_{2n+2}.

Alkenes: A group of aliphatic hydrocarbons that contain one or more double carbon-carbon bonds, where the carbon atoms share two pairs of electrons. The general molecular formula for alkanes is C_nH_{2n}.

Alkyl group: A hydrocarbon substituent that includes any branch that is a hydrocarbon with only single bonds. All alkyl groups use the prefixes for the number of carbons followed by the ending *–yl*.

Alkynes: Hydrocarbons that contain one or more triple carbon-carbon bonds, where the carbon atoms share three pairs of electrons. The general molecular formula for alkynes is C_nH_{2n-2}.

Allotrope: In each different allotrope, an element's atoms are bonded together in a different manner.

Alloy: A solution made from two metals, formed by mixing molten metals and allowing them to solidify.

Amalgam: A mixture of mercury and an alloy of tin, silver, and copper.

Amontons, Guillaume: Developed a thermometer based on the increasing volume of a gas (instead of a liquid) with an increase in temperature.

Anion: A negative ion formed when an atom gains one or more electrons.

***(aq)*:** Identifies that the substance is aqueous, or dissolved in water.

Aqueous solution: Water that contains dissolved substances.

Aromatic hydrocarbons/arenes: Organic compounds that contain a benzene ring.

Atmosphere (atm): A unit of pressure. One atmosphere is equal to 760 mmHg, or 101.325 kPa.

Atomic mass units (amu, u or μ): Defined as $\frac{1}{12}$ the mass of a carbon-12 atom.

Atomic (Z) number: The number of protons in the nucleus of an atom of an element.

Avogadro, Amadeo: Published Avogadro's Hypothesis is 1811, which stated that a sample of any gas at the same temperature and pressure will contain the same number of particles.

Avogadro's Hypothesis/Avogadro's Principle: Published in 1811, it stated that equal volumes of gases at the same temperature and pressure will contain the same number of particles.

Average atomic mass: The weighted average of the relative abundances of each isotope of an element.

Average kinetic energy: The amount of kinetic energy found midway between the extremes of the various energies possessed by gas particles in motion in a sample of matter.

Avogadro's Number (N_A): 6.02×10^{23} particles/mole.

Balanced chemical equation: A chemical equation that has an equal number of atoms for each element on both the reactant and product side of the equation.

Barometer: An instrument used to determine atmospheric pressure.

Binary ionic compounds: Usually contain one positively charged metal ion combined with one negatively charged non-metal ion.

Bitumen: Also known as asphalt or tar, it is a black, oily, viscous material that is a naturally occurring organic byproduct of decomposed organic materials.

Boiling: Occurs when vaporization takes place throughout the liquid and the temperature of the liquid remains constant.

Boiling point: The temperature at which the vapour pressure of a substance is equal to the external or atmospheric pressure.

Boiling point elevation: When solute is added to a solution, the solute particles lower the vapour pressure. More heat energy is then required to bring the temperature to where the vapour pressure of the liquid is equal to the air pressure (i.e., to the boiling point). This results in a higher boiling point compared to that of the solvent itself.

Boyle's Law: States that the pressure of a gas is inversely proportional to its volume when temperature and the amount of gas are held constant.

Boyle's Law equation: $P_1V_1 = P_2V_2$.

Buckminsterfullerene/buckyball/fullerene: An allotrope of carbon consisting of interlocking hexagons and pentagons.

Carboxylic acids: Contain one or more carboxyl groups. The general form for a carboxylic acid is *R–COOH*, where *R* represents a hydrocarbon chain.

Catalytic cracking: A process in which a catalyst (such as a sample of small pellets of silica) is added to petroleum in the absence of oxygen to make substances that are used in the making of gasoline and starting molecules for synthetic petroleum products.

Cations: Positively charged ions formed when an atom loses one or more electrons.

Celsius, Anders: Invented the centigrade (Celsius) temperature scale, using the freezing point of water as 0°C and the boiling point of water as 100°C.

Charles, Jacques: Investigated the expansion rates of different gases due to temperature changes.

Charles' Law: States that as temperature increases, so does the volume of a sample of gas when the pressure and amount of gas are held constant.

Charles' Law equation: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Chemical bond: The force of attraction holding atoms together to make compounds.

Chemical equation: Indicates the substances reacting and the substances produced in a chemical reaction.

Chemical formula: A shorthand method to represent compounds that uses the elements' symbols and subscripts.

Chemical properties: Describe how a substance reacts with other substances (e.g., rusting, creating gas bubbles).

Coefficients: Indicate the ratio in which the substances combine or are produced in a chemical reaction.

Colligative property: A property that depends on the number of particles of solute, and not upon their identity. There are three colligative properties that are important to understanding solutions—vapour pressure lowering, freezing point depression and boiling point elevation.

Colloid: A mixture of particles that are smaller than those in suspensions, but larger than the particles in solutions. The particles in colloids can be evenly distributed throughout a dispersion medium. They have a characteristic cloudy or milky appearance.

Combined Gas Law: Describes the relationship between the pressure, volume and temperature of a contained gas.

Combined Gas Law equation: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Combustion: The process in which a substance reacts with oxygen. Complete combustion occurs when there is sufficient oxygen present, creating carbon dioxide and water vapour. Incomplete combustion occurs when there is insufficient oxygen, creating carbon dioxide, carbon monoxide, and water vapour.

Combustion reaction: A chemical change where oxygen reacts with another element or compound, often resulting in the production of heat or light energy. The general equation for a combustion reaction is:

$$C_xH_y + \left(\frac{x+y}{4}\right)O_2 \rightarrow xCO_2 + \left(\frac{y}{2}\right)H_2O$$

Compressibility: The property of being able to occupy less space.

Concentrated solution: A solution that has a lot of solute as compared to the amount of solvent present.

Concentration: The amount of solute per unit of solution (solute plus solvent).

Condensation/Liquefaction: The conversion of a gas to a liquid. It is an exothermic process.

Condensation polymers: Formed by the reaction of two or more different monomers. During this polymerization process, small molecules such as water, ammonia, or hydrogen chloride are produced.

Condensed structural formulas: Formulas that save space by showing the atoms in an organic compound, but not their bonds.

Covalent bonds: A chemical bond formed as a result of the sharing of electron pairs.

Crude oil: The name given to the unprocessed petroleum that is extracted from the ground. The refining of crude oil involves several steps.

Cyclic hydrocarbons: Compounds that contain a hydrocarbon ring.

Dalton (Da): An alternate name for atomic mass unit (amu, u or μ) used in biochemistry and molecular biology.

Dalton, John: Stated that in a mixture of gases the total pressure is equal to the sum of the pressure of each gas, as if it were in a container alone.

Dalton's Law of Partial Pressures: States that the partial pressure of a gas in a mixture of gases is equal to the pressure the gas would exert if it occupied the same volume alone.

Decomposition reactions: Have only one reactant producing two or more simpler substances, usually requiring the input of energy in the form of heat, electricity, or light. The general form of a decomposition reaction is
$AB \rightarrow A + B$.

Decompression sickness/the "bends": Severe joint pain, vomiting, and dizziness caused when divers ascend too quickly and bubbles of released nitrogen gas block small blood vessels and limit oxygen uptake by cells.

Dehydrogenation: The removal of hydrogen atoms to convert an alkane to an alkene or an alkene to an alkyne.

Density: A ratio that compares the mass of an object to its volume.

Deposition: An exothermic process involving the direct conversion to a solid from a vapour without forming a liquid.

Diamond: An allotrope of carbon in which the carbon atoms are bonded together in a tetrahedral lattice arrangement.

Diatomic molecules: Elements that exist as pairs of molecules joined covalently.

Diffusion: The movement of one substance through another.

Dilute solution: A solution that has little solute as compared to the amount of solvent present.

Dipolar molecule/dipole: A molecule that has two poles (electrically charged regions).

Dissociate/dissociation: The act of a compound breaking up into positive and negative ions.

Double replacement/double displacement reactions: Occur when the positive ions of two compounds change positions. These types of reactions generally produce a precipitate, a gas, or a molecular compound, such as water. The general form of the reaction is $AX + BY \rightarrow BX + AY$.

Dynamic equilibrium: A state of balance in which the rate of evaporation is equal to the rate of condensation in a closed container.

Elastic collision: A collision in which there is no loss of energy.

Electrolyte: A solution that conducts an electric current.

Electron affinity: A measure of an atom's ability to gain electrons.

Electronegativity: The attraction that an atom has for the shared electrons in a covalent bond.

Electrons: Negatively charged particles that surround the nucleus of an atom.

Empirical formula: The formula of a compound having the smallest whole number ratio of the moles of each element. It must be obtained from experimental data.

Emulsifying agent: Added to emulsions to keep the suspension of liquids from separating from each other upon standing.

Endothermic: Describes a process or reaction that requires that energy be added to the system.

Energy ratio: The ratio of heat to the coefficient of one other substance in a chemical equation.

Esterification: The process of forming an ester from the reaction between an organic acid and an alcohol in the presence of a catalyst, which is usually concentrated sulfuric acid.

Evaporation: The conversion of a liquid to a gas that occurs specifically on the surface of a liquid.

Excess reactants: The remaining, or leftover, reactants after a reaction has taken place.

Exothermic: Describes a process or reaction in which more energy is released by the system than is required to break bonds.

Extrapolation: The determination of values that fall beyond measured points on a graph by extending the line of best fit.

Fahrenheit, Daniel: Invented the mercury thermometer and the temperature scale that is named after him. He proposed 0 to be the coldest temperature in Western Europe and 100 to be the highest temperature. This made the freezing point of water 32°F and the boiling point of water 212°F.

Fertilizer: Any compound that contains one or more chemical elements, organic or inorganic, natural or synthetic.

Formula mass: The sum of the atomic masses of all of the atoms in one molecule or particle of that substance, in amu's.

Fractional distillation/fractionation: The process used to separate organic compounds based on their boiling points.

Freezing/Solidifcation: The conversion of a liquid to a solid through the removal of energy from a sample of particles so that the particles slow down and settle closer together due to stronger intermolecular forces of attraction. It is an exothermic process.

Freezing point: The temperature at which a substance freezes. It is unique for every substance.

Freezing point depression: When solute is added to a solvent, the solute lowers the vapour pressure of the solvent. As a result, temperature must be lowered below the freezing point of the solvent in order for the vapour pressure of the solid and the vapour pressure of the liquid to be at equilibrium. The greater the amount of solute, the greater the freezing point depression (i.e., the lower the freezing point).

Functional groups: Small units within an organic molecule that are responsible for most of the chemical and physical properties of that molecule.

***(g)*:** Gas or vapour.

Galilei, Galileo: His suction pump allowed him to determine that the limit to which the height of water can be raised is 32 feet (11 m).

Gas: The state of matter distinguished from solid and liquid states by a relatively low density and viscosity.

Gas pressure: The force of gas particles in their container acting on each square centimetre of area.

Gay-Lussac, Joseph Louis: Proposed the law of combining volumes, and the direct relationship between the pressure and temperature of a gas.

Gay-Lussac's Law: States that as temperature increases, so does the pressure of a sample of gas when the volume and amount of gas are held constant.

Gay-Lussac's Law equation: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Graphite: An allotrope of carbon in which the carbon atoms are bonded together in sheets of a hexagonal lattice.

Greenhouse gas: A gas, such as methane or carbon dioxide, that traps heat energy within the Earth's atmosphere.

Heat of hydration/energy of hydration: The energy that is released when solute ions move between solvent particles, and the forces of attraction between solute and solvent particles take hold so that the particles "snap" back and move closer.

Heat of solution: The total heat change in the dissolving process.

Homogeneous: Uniform throughout a sample of material, such that every part of the material is exactly the same as any other part in terms of appearance and phase.

Huygens, Christiaan: Developed the manometer to study the elastic forces in gases.

Hydrated: When water surrounds individual molecules or ions.

Hydrates: Solids that have water molecules tightly associated with their solid crystal.

Hydrocarbon: The simplest form of an organic compound, composed of hydrogen and carbon. Most hydrocarbons are extracted from petroleum.

Hydrocracking: Catalytic cracking in the presence of hydrogen, whereby the extra hydrogen saturates, or hydrogenates, the chemical bonds of the cracked hydrocarbons. Hydrocracking is also a treating process, because the hydrogen combines with contaminants such as sulfur and nitrogen, allowing them to be removed.

Hydrogenation: The addition of hydrogen atoms to convert an alkyne to an alkene or an alkene to an alkane.

Hydrogen bonds: The intermolecular interaction among water molecules resulting from the partial positive end of the water molecule being attracted to the negative end of another molecule. Hydrogen bonding is stronger than the normal attractive force between neighbouring molecules.

Immiscible: The term used to describe two liquids that do not dissolve in any proportion.

Initiator: A substance (usually a peroxide) that is used to start the polymerization process by producing a free radical.

Insoluble: The term used when a substance does not dissolve in a solvent.

Intermolecular forces: The forces of attraction between particles.

Interpolation: The determination of values that fall between measured points on a line of best fit.

Ion: A charged particle that forms when a neutral atom gains or loses electrons.

Ionic bond: Formed when a negatively charged ion is attracted to a positively charged ion.

Ionic compounds: Formed when two or more oppositely charged ions are attracted to each other, usually made of a metal and a non-metal.

Ionization energy: A measure of an atom's ability to lose electrons.

Isotopes: Atoms that have the same number of protons but differ in their number of neutrons.

Kilopascal: One kilopascal is equal to 1000 Pascals.

Kinetic energy: The energy of objects in motion.

Kinetic Molecular Theory: Describes the size, motion, and energy of particles.

***(l)*:** Liquid.

Lattice energy: The amount of energy required to separate the molecules or ions from each other in a solid crystal.

Law of Conservation of Mass: States that mass cannot be created or destroyed.

Limiting reactant/limiting reagent/limiting factor: The reactant that becomes completely consumed in a chemical reaction.

Liquid: A fluid matter having no fixed shape, but a fixed volume.

Manometer: An open or closed device used to measure the vapour pressure exerted by a vapour over a liquid.

Mass (A) number: The sum of the number of protons and neutrons found in the nucleus of an atom.

Matter: Anything that takes up space (has a volume) and has a mass.

Melting/Fusion: An endothermic process, in which energy is absorbed to overcome the intermolecular forces of attraction between solid particles so that a substance's particles can move more freely in the liquid state.

Melting point: The temperature at which a substance melts or fuses. It is unique for every substance.

Meniscus: The curvature of the surface of the water.

Millibar: A meteorological unit of atmospheric pressure (1013 mbars = 1.013 bar = 1 atmosphere).

Millimetres of mercury (mmHg): A unit of pressure traditionally used with barometers (760 mmHg = 1 atmosphere).

Miscible: The term used to describe two liquids that dissolve in each other in any proportion.

Mixture: Contains more than one kind of particle.

Molar coefficients: The number of moles that combine or are produced as indicated by the coefficients in the balanced equation.

Molar mass/molecular weight (MM): The mass of one mole of a substance. Its units are in grams per mole (g/mol).

Molar ratio/mole ratio: The ratio of coefficients for two substances in a chemical equation.

Molar volume: The volume occupied by one mole of a gas. At STP, this value is 22.4 L/mole; at SATP, the value is 24.4 L/mol.

Mole (mol): The unit that relates the number of particles in a sample to its mass. It is the number of particles in 12.0000 g of carbon atoms containing 6 protons and 6 neutrons each. In one mole of carbon, there are 6.02×10^{23} atoms.

Molecular/covalent compound: Neutral compounds made of atoms joined covalently.

Molecular formula: Contains the actual number of atoms of each element in one unit of that compound. The subscripts of the molecular formula are always a whole number multiple of the subscripts of the empirical formula.

Molecular formula of an organic compound: Shows the kind of atoms and the number of each kind of atom in the compound.

Molecular mass: The sum of the atomic masses of all of the atoms in a molecule.

Monomers: The basic building blocks of polymers.

Natural gas: A mixture of several naturally occurring gases in the atmosphere, consisting of 80% methane, 10% ethane, 4% propane, 2% butane, and 4% nitrogen, helium, and other trace gases.

Neutrons: Neutral particles found within the nucleus of an atom.

Non-electrolyte: A solution that does not conduct an electric current.

Nonpolar bonds: Bonds formed between atoms of similar electronegativity (they differ by less than 0.2).

Nonpolar molecules: Molecules held together by nonpolar bonds.

Normal/straight-chain/unbranched hydrocarbons: Alkanes in which the carbon atoms form long chains.

Organic: Substances originally thought to have come directly from living organisms (that is, plant and animal in nature).

Organic chemistry: The study of the structure, composition, properties, and reactions of organic compounds. Today, organic chemistry includes the study of ALL carbon-based compounds, regardless of their origin.

Particle: A generic term meaning atoms, molecules, ions, or formula units.

Parts per billion (ppb): Refers to the number of grams of solute for every one billion grams of solution, or the number of particles of solute for every one billion particles of solvent.

Parts per million (ppm): Refers to the number of grams of solute for every one million grams of solution, or the number of particles of solute for every one million particles of solvent.

Pascal (Pa): One Newton of force per square metre. One Pascal is equal to the pressure exerted by a stamp on an envelope.

Pascal, Blaise: Discovered that the pressure of the atmosphere increases as you move down a mountain.

Percentage volume by volume (%v/v): The volume (in mL) of solute for every 100 mL of solution (volume/volume) or

$$\%\text{v/v} = \frac{\text{mL of solute}}{\text{100 mL solution}} \times 100\%$$

Percentage weight by volume (%w/v): The mass of solute for every 100 mL of solution (weight/volume) or

$$\%\text{w/v} = \frac{\text{grams of solute}}{\text{100 mL solution}} \times 100\%$$

Percentage weight by weight (% w/w): The grams of solute in 100 g of solution (weight/weight) or

$$\%\text{w/w} = \frac{\text{grams of solute}}{\text{100 g solution}} \times 100\%$$

Percent composition: Determined by dividing the mass of the element in the compound by the total mass of the compound and multiplying by one hundred.

Percent yield: The ratio of the actual yield to the theoretical yield expressed as a percentage.

Petroleum: Derived from the Latin roots 'petra' meaning rock and 'oleum' for oil.

Phase: The mixture of states of matter that coexist as physically distinct parts of a mixture.

Phenyl group: When the benzene ring is a substituent of a hydrocarbon chain.

Photosynthesis: A reaction in which blue-green algae, some bacteria, and most green plants convert sunlight, carbon dioxide, and water vapour into carbohydrates and oxygen.

Physical properties: Properties that we can observe without chemically changing a substance (e.g., hardness, colour).

Plasma: A gaseous mixture of positive ions and electrons. It is the fourth state of matter distinct from solids, liquids, and gases, which is present in stars and fusion reactors.

Polar bonds: Bonds formed between atoms of differing electronegativities.

Polar molecule: A molecule held together by polar bonds.

Polyatomic ions: Ions that are composed of several atoms joined covalently.

Polymer: A very large molecule made of many smaller repeating units called monomers.

Pounds per square inch (psi): An Imperial unit of pressure often encountered when filling car and bicycle tires. One kilopascal is equal to 0.145 psi.

Precipitate: Means to return to the undissolved state from the solution.

Products: The resulting substance or substances in a chemical reaction.

Property: A characteristic that we can use to help us identify a person, place, or thing.

Protons: Positively charged particles found within the nucleus of an atom.

Pure substance: Contains only one kind of particle.

Qualitative properties: Descriptive information based on the observation of physical characteristics (e.g., orange flame, yellow gold, lustrous copper).

Quantitative properties: Numerical information (e.g., 10 grams of magnesium burned; the density of water is 1.0 g / cm^3).

Reactants: The substance or substances that react together in a chemical reaction.

Relative natural abundance: The fraction of each isotope found in an average sample of the element.

Resonance: When two or more equally correct structures can be drawn for one molecule.

***(s)/(c)*:** Solid or crystalline.

SATP (Standard Ambient Temperature and Pressure): Refers to a temperature of 25ºC and either 101.3kPa or 1 atm of pressure.

Saturated: A solution that contains the maximum amount of solute for that amount of solvent at a particular temperature. It is generally very stable.

Saturated organic compound: Those compounds whose carbon-carbon bonds are all single bonds. This means that each carbon is bound to four atoms, the maximum number possible.

Scientific notation: A method of writing values in terms of a decimal number greater than 1 and less than 10 multiplied by a power of 10.

Significant digits (a.k.a. significant figures): All digits from a measured value that are known (or certain), plus the last digit that is estimated.

Single replacement/single displacement reactions: Reactions involving a compound and an element. The element replaces (or displaces) one element in the compound to produce a new element and a new compound. In these reactions, metals replace metals (cations replace cations) and non-metals replace non-metals (anions replace anions). The general form of the single replacement reactions is $A + BX \rightarrow B + AX$ or $Y + AX \rightarrow X + AY$.

Skeleton equation: A reaction that includes the formulas of each reactant and product.

Solid: A state of matter in which there is a definite shape and volume, neither liquid nor gaseous.

Solubility: The maximum amount of solute that can dissolve in a certain amount of solvent, or the amount of solute needed to make a saturated solution under certain conditions of temperature and pressure.

Solubility curve: Each point on the curve represents a saturated solution of a solute in water at a different temperature. The area *below* the curve represents quantities that produce an unsaturated solution at that temperature. The points *above* the curve tend to indicate a supersaturated solution.

Soluble: The term used when a substance is able to dissolve in a solvent.

Solute: The dissolved particles in a solution. Usually present in a smaller amount.

Solutions: Homogeneous mixtures that are composed of two or more substances that are evenly distributed throughout a single phase.

Solvation: The process by which positive and negative ions become surrounded by solvent molecules.

Solvent: The dissolving medium in a solution. Usually present in a larger amount.

Sphygmomanometer: The tool used to measure the pressure of blood on the walls of arteries when your blood pressure is taken.

States of matter: The distinct forms taken on by different phases of matter.

Stock naming system: Uses Roman numerals following the metal ion's name to indicate an ion's charge.

Stock solutions: A more concentrated solution. Using a stock solution, you can make a solution of any lower concentration, simply by adding more solvent.

Stoichiometry: The use of ratios to determine the quantities of reactants used and products produced in a chemical reaction.

STP (Standard Temperature and Pressure): $0^{o}C$ and either 101.3kPa or 1 atm of pressure.

Structural formula of an organic compound: Shows the arrangement of the atoms found in the molecular formula.

Structural isomers: Hydrocarbons with the same molecular formula but different arrangements of the atoms.

Sublimation: An endothermic process in which a solid changes directly to a gas without passing through the liquid state.

Substituent: An atom or group of atoms that can take the place of a hydrogen atom on a parent hydrocarbon molecule.

Supersaturated: A solution that contains more solute than the solvent can normally hold at that temperature. It is very unstable.

Suspension: Contains large particles that settle out upon standing. You can use a filter to separate the components of a suspension, whereas you cannot accomplish this with a solution.

Synthesis reactions: Involve the reaction of two simple substances to produce a single, more complex substance. The general form of a synthesis reaction is $A + B \rightarrow AB$.

Synthetically produced hydrocarbons: Constructed by starting with a petrochemical compound and adding to it to create longer hydrocarbon chains. Many synthetic products are more stable at higher temperatures, and are very insoluble, making them excellent lubricants.

Tar sands/oil sands: A combination of clay, sand, water, and bitumen.

Theoretical yield: The maximum amount of product you could possibly form in a chemical reaction.

Thermal cracking: A process that uses heat to break down the residue from vacuum distillation. The lighter elements produced from this process can be made into distillate fuels and petrol.

Thermal pollution: Occurs when an industrial plant draws water from a nearby lake, uses the water, and then returns the water to the lake at a higher temperature. The temperature of the entire lake then increases, decreasing the oxygen solubility in the lake water.

Thomson, William: Created the Kelvin scale where absolute zero (–273°C) was the lowest temperature possible. He further reasoned that at this temperature all molecular motion would cease, the kinetic energy would be zero, and the volume of a gas, hypothetically, would also be zero.

Torricelli, Evangelista: He developed the barometer, and used it to determine that the limit to the height to which Galileo's pump could draw water was due to atmospheric pressure.

Unsaturated: A solution that can hold more solute at that temperature.

Unsaturated hydrocarbons: Organic compounds that contain double and triple C—C bonds. They possess fewer than the maximum number of hydrogens in their structure.

Vacuum distillation: Allows heavy hydrocarbons with boiling points of 450°C and higher to be separated.

Vaporization: The conversion of a liquid to a gas through the addition of energy to a sample of particles. As a result of the additional energy, the particles move faster, overcome their intermolecular forces of attraction, and enter the gas phase. It is an endothermic process.

Vapour pressure, P_{vap}: The pressure exerted by a gas on the sides of its container when the liquid and gaseous phases of a substance are at dynamic equilibrium.

Viscosity: The measure of resistance of a liquid to flow, dependent on the strength of the intermolecular forces of attraction, particle size, shape, and temperature.

Vital force: A theory proposed by Jons Jakob Berzelius, which assumed that organic compounds could only be formed in living cells, and that it was impossible to prepare them in laboratories.

Volatile: Substances having weak intermolecular forces of attraction that are easily overcome. They disperse into the air quickly.

von Guericke, Otto: Made a pump that could create a vacuum so strong that a team of 16 horses could not pull two metal hemispheres apart.

Word equation: The description of a chemical reaction using words.

Appendix B: Web Resources

This appendix is a collection of web resources that are referenced in the course. All links were active as of August 2010, but you may have difficulty accessing some of the sites with the URL provided. If so, try using the author and title information to find the resource with a Google search.

Module 1: Physical Properties of Matter

Lesson 2: Kinetic Molecular Theory

Goodman, J.M. *Can Crush*. 2007
www.youtube.com/watch?v=1Efy36yxdUc

Harcourt School Publishers. *States of Matter*. 2010.
www.harcourtschool.com/activity/states_of_matter/

Module 2: Gases and the Atmosphere

Lesson 1: Gases in the Atmosphere

Anderson, Larry S. *Format for Update of Information Technology Plans*. 1996.
www.nctp.com/html/tech_plan_format.cfm

Lesson 3: Gas Pressure and Volume, Part A

Krampf, Robert. *Cartesian Diver*. 2008.
www.metacafe.com/watch/877495/cartesian_diver/

———. *In Space Without a Spacesuit*. 2008.
www.metacafe.com/watch/898336/in_space_without_a_spacesuit

Lesson 6: Gas Temperature and Volume, Part A

Douglas, Robert. *Fire Syringe* Thermodynamics. 2008.
http://teachertube.com/viewVideo.php?video_id=55314&title=Fire_Syringe_Thermodynamics

Lesson 8: Gas Pressure and Temperature , Part A

Kentchemistry.com. The Dancing Penny. 2009.
www.metacafe.com/watch/1001831/simple_science_the_dancing_penny/

Lesson 11: Applications of the Gas Laws

McGraw-Hill Higher Education. *Diving and Gases*. 2009
www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/gam2s3_2.rm

Module 3: Chemical Reactions

Lesson 1: Atomic Mass and Isotopes

U.S. Geological Survey. *Resources on Isotopes*. 2004.
http://wwwrcamnl.wr.usgs.gov/isoig/period/h_iig.html

U.S. Environmental Protection Agency. Tritium. 2009.
www.epa.gov/radiation/radionuclides/tritium.html

Lesson 4: Balancing Chemical Equations

Earley, Clarke. *Balancing Chemical Equations*. 2008.
www.personal.kent.edu/~cearley/ChemWrld/balance/balance.htm

Bishop, Mark. *An Introduction to Chemistry*. "Balancing Chemical Equations." 2010
http://preparatorychemistry.com/Bishop_Balancing_Equations_help.htm

Fun Based Learning. *Chemistry Games*. 2006.
www.funbasedlearning.com/chemistry/chembalancer/default. htm
www.funbasedlearning.com/chemistry/chembalancer2/default.htm
www.funbasedlearning.com/chemistry/chembalancer3/default.htm

Kandarsjah, Kosasihis. *Five Major Chemical Reactions*. 2008.
www.youtube.com/watch?v=tE4668aarck

Lesson 5: Chemical Reactions

VirginiaTech. *Practice Predicting Products of Reactions*. n.d.
www.files.chem.vt.edu/RVGS/ACT/notes/Practice_Predicting.html

Module 4: Stoichiometry

Lesson 4: Limiting Reactant

W.W. Norton. *Limiting Reactant*. 2010.
www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_03&folder=limiting_reactants

Petrucci, Ralph H., William S. Harwood, and Geoffrey Herring. *General Chemistry: Principles and Modern Applications.* "Chapter 4: Chemical Reactions." 2002.
http://cwx.prenhall.com/petrucci/medialib/media_portfolio/04.html

Chang, Raymond. *Essential Chemistry, 2/e*. 2000.
www.mhhe.com/physsci/chemistry/essentialchemistry/flash/flash.mhtml

Module 5: Solutions

Lesson 3: The Solution Process

Northland College. *A Quick Look at How Ionic Compounds Dissolve*. 2010.
http://programs.northlandcollege.edu/biology/Biology1111/animations/dissolve.html

Lesson 4: Dissolving Polar and Non-Polar Substances

Spangler, Steve. *Eco-Friendly Packing Peanuts Demonstration*. 2008.
www.youtube.com/watch?v=PNz198hMboQ

Lesson 8: Freezing Point Depression and Boiling Point Elevation

Kentchemistry.com. *Instantly Freeze Soda Experiment.* 2007.
www.metacafe.com/watch/382767/instantly_freeze_soda_experiment/

Lesson 10: Preparing a Solution

Devarela, M. Nunez. *Preparation of a Standard Solution.* 2007.
www.youtube.com/watch?v=XMtm4hVCGWg

Lesson 11: Dilutions

Lowlevelpanic999. *Dilutions*. 2010.
www.youtube.com/watch?v=3EOiCrtvUUM

Lesson 12: Concentration Applications

Centers for Disease Control and Prevention. *Facts about Chlorine*. 2006. www.bt.cdc.gov/agent/chlorine/basics/facts.asp

Government of Manitoba. *Water Stewardship Branch*. “Office of Drinking Water.” 2010.
www.gov.mb.ca/waterstewardship/odw/

Module 6: Organic Chemistry

Lesson 11: Chemistry and Industry

Discovery Channel. *How Stuff Works: Rubber and Spaghetti*. 2009. http://dsc.discovery.com/videos/howstuffworks-rubber-and-spaghetti.html

Macrogalleria. Polymers. n.d.
http://pslc.ws/macrog/index.htm

Spangler, Steve. *Oil Spill Absorbing Polymer*. 2010.
www.stevespanglerscience.com/experiment/00000108

———. *Baby Diaper Secret*. 2010.
www.stevespanglerscience.com/experiment/00000064

———. *Bubble Gum Science*. 2010.
www.stevespanglerscience.com/experiment/00000133

———. *GAK – Elmer’s Glue Borax Recipe*. 2010.
www.stevespanglerscience.com/experiment/00000039

Appendix C: Names, Formulas, and Charges of Common Ions

Positive Ions (Cations)

Name	Symbol	Name	Symbol
aluminum	Al^{3+}	magnesium	Mg^{2+}
ammonium	NH_4^+	**manganese(II)**	**Mn^{2+}**
barium	Ba^{2+}	manganese(IV)	Mn^{4+}
cadmium	Cd^{2+}	mercury(I)	Hg_2^{2+}
calcium	Ca^{2+}	mercury(II)	Hg^{2+}
chromium(II)	Cr^{2+}	**nickel(II)**	**Ni^{2+}**
chromium(III)	**Cr^{3+}**	nickel(III)	Ni^{3+}
copper(I)	Cu^+	potassium	K^+
copper(II)	**Cu^{2+}**	silver	Ag^+
hydrogen	H^+	sodium	Na^+
iron(II)	Fe^{2+}	strontium	Sr^{2+}
iron(III)	**Fe^{3+}**	tin(II)	Sn^{2+}
lead(II)	**Pb^{2+}**	**tin(IV)**	**Sn^{4+}**
lead(IV)	Pb^{4+}	zinc	Zn^{2+}
lithium	Li^+		

continued

Negative Ions (Anions)

Name	Symbol	Name	Symbol
acetate	$C_2H_3O_2^-$ (CH_3COO^-)	nitrate	NO_3^-
azide	N_3^-	nitride	N^{3-}
bromide	Br^-	nitrite	NO_2^-
bromate	BrO_3^-	oxalate	$C_2O_4^{2-}$
carbonate	CO_3^{2-}	hydrogen oxalate	$HC_2O_4^-$
hydride	H^-	oxide	O^{2-}
hydrogen carbonate or bicarbonate	HCO_3^-	perchlorate	ClO_4^-
chlorate	ClO_3^-	permanganate	MnO_4^-
chloride	Cl^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}
chromate	CrO_4^{2-}	dihydrogen phosphate	$H_2PO_4^-$
citrate	$C_6H_5O_7^{3-}$	silicate	SiO_3^{2-}
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfate	HSO_4^-
fluoride	F^-	sulfide	S^{2-}
hydroxide	OH^-	hydrogen sulfide	HS^-
hypochlorite	ClO^-	sulfite	SO_3^{2-}
iodide	I^-	hydrogen sulfite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Appendix D: Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4^+ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

(continued)

Anions (Negative Ions)

1⁻ charge		1⁻ charge		2⁻ charge	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^-	Bromate	OH^-	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^-	Bromide	IO_3^-	Iodate	O^{2-}	Oxide
ClO_3^-	Chlorate	I^-	Iodide	O_2^{2-}	Peroxide
Cl^-	Chloride	NO_3^-	Nitrate	SO_4^{2-}	Sulfate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulfide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulfite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocyanate	**3⁻ charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Appendix E: Periodic Table of Elements

Atomic Number → 19
Symbol → **K**
Name → Potassium
Relative Atomic Mass → 39.1

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Ununhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Appendix F: Alphabetical Listing of the Elements and Their Atomic Masses

Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Actinium	(227)	Gold	197.0	Praseodymium	140.9
Aluminum	27.0	Hafnium	178.5	Promethium	(145)
Americium	(243)	Hassium	(265)	Protactinum	(231)
Antimony	121.7	Helium	4.0	Radium	(226)
Argon	39.9	Holmium	164.9	Radon	(222)
Arsenic	74.9	Hydrogen	1.0	Rhenium	186.2
Astatine	(210)	Indium	114.8	Rhodium	102.9
Barium	137.3	Iodine	126.9	Rubidium	85.5
Berkelium	(247)	Iridium	192.2	Ruthenium	101.1
Beryllium	9.0	Iron	55.8	Rutherfordium	(261)
Bismuth	209.0	Krypton	83.8	Samarium	150.4
Bohrium	(264)	Lanthanum	138.9	Scandium	45.0
Boron	10.8	Lawrencium	(257)	Seaborgium	(263)
Bromine	79.9	Lead	207.2	Selenium	79.0
Cadmium	112.4	Lithium	6.9	Silicon	28.1
Calcium	40.1	Lutetium	175.0	Silver	107.9
Californium	(251)	Magnesium	24.3	Sodium	23.0
Carbon	12.0	Manganese	54.9	Strontium	87.6
Cerium	140.1	Meitnerium	(266)	Sulfur	32.1
Cesium	132.9	Mendelevium	(256)	Tantalum	180.9
Chlorine	35.5	Mercury	200.6	Technetium	(98)
Chromium	52.0	Molybdenum	95.9	Tellurium	127.6
Cobalt	58.9	Neodymium	144.2	Terbium	158.9
Copernicium	(277)	Neon	20.2	Thallium	204.4
Copper	63.5	Neptunium	(237)	Thorium	232.0
Curium	(247)	Nickel	58.7	Thulium	168.9
Dubnium	(262)	Niobium	92.9	Tin	118.7
Dysprosium	162.5	Nitrogen	14.0	Titanium	47.9
Einsteinium	(254)	Nobelium	(259)	Tungsten	183.8
Erbium	167.3	Osmium	190.2	Uranium	238.0
Europium	152.0	Oxygen	16.0	Vanadium	50.9
Fermium	(257)	Palladium	106.4	Xenon	131.3
Fluorine	19.0	Phosphorus	31.0	Ytterbium	173.0
Francium	(223)	Platinum	195.1	Yttrium	88.9
Gadolinium	157.2	Plutonium	(244)	Zinc	65.4
Gallium	69.7	Polonium	(209)	Zirconium	91.2
Germanium	72.6	Potassium	39.1		

Appendix G: Table of Electronegativities

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 **Lanthanide Series**	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 **Actinide Series**	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Appendix H: List of SLOs Found in Each Lesson

This appendix has been developed for classroom teachers who are interested in knowing which Specific Learning Outcomes are taught in each lesson in this Independent Study course.

Module 1: Physical Properties of Matter

Lesson 1: States of Matter

C11-1-01 Describe the properties of gases, liquids, solids, and plasma.
Include: density, compressibility, diffusion

Lesson 2: Kinetic Molecular Theory

C11-1-02 Use the Kinetic Molecular Theory to explain properties of gases.
Include: random motion, intermolecular forces, elastic collisions, average kinetic energy, temperature

C11-1-03 Explain the properties of liquids and solids using the Kinetic Molecular Theory.

Lesson 3: Phase Changes

C11-1-04 Explain the process of melting, solidification, sublimation, and deposition in terms of the Kinetic Molecular Theory.
Include: freezing point, exothermic, endothermic

C11-1-05 Use the Kinetic Molecular Theory to explain the processes of evaporation and condensation.
Include: intermolecular forces, random motion, volatility, dynamic equilibrium

Lesson 4: Vapour Presuure

C11-1-06 Operationally define vapour pressure in terms of observable and measurable properties.

C11-1-07 Operationally define normal boiling point temperature in terms of vapour pressure.

Lesson 5: Graphing Vapour Pressure

C11-1-08B Interpolate and extrapolate the vapour pressure and boiling temperature of various substances from pressure versus temperature graphs.

Module 2: Gases and the Atmosphere

Lesson 1: Gases in the Atmosphere

C11-2-01 Identify the abundances of the naturally occurring gases in the atmosphere and examine how these abundances have changed over geologic time.
Include: oxygenation of Earth's atmosphere, the role of biota in oxygenation, changes in carbon dioxide content over time

C11-2-02 Research Canadian and global initiatives to improve air quality.

Lesson 2: Gas Pressure

C11-2-03 Examine the historical development of the measurement of pressure.
Examples: the contributions of Galileo Galilei, Evangelista Torricelle, Otto von Guericke, Blaise Pascal, Christiaan Huygens, John Dalton, Joseph Louis Gay-Lussac, Amadeo Avogadro

C11-2-04 Describe the various units used to measure pressure.
Include: atmospheres (atm), kilopascals (kPa), millimetres of mercury (mmHg), millibars (mb)

Lesson 3: Gas Pressure and Volume, Part A

C11-2-05A Experiment to develop the relationship between the pressure and volume of a gas using visual and graphical representations.
Include: historical contributions of Robert Boyle

Lesson 4: Working with Significant Figures

Understand the importance of expressing calculation answers to the correct number of significant figures.

Perform calculations that result in the correct number of significant figures.

Lesson 5: Gas Pressure and Volume, Part B

C11-2-05B Experiment to develop the relationship between the pressure and volume of a gas using numeric representation.
Include: historical contributions of Robert Boyle

Lesson 6: Gas Temperature and Volume, Part A

C11-2-06A Experiment to develop the relationship between the volume and temperature of a gas using visual and graphical representations.
Include: historical contributions of Jacques Charles, the determination of absolute zero, the Kelvin temperature scale

Lesson 7: Gas Temperature and Volume, Part B

C11-2-06B Experiment to develop the relationship between the volume and temperature of a gas using numeric representation.
Include: historical contributions of Jacques Charles, the determination of absolute zero, the Kelvin temperature scale

Lesson 8: Gas Pressure and Temperature, Part A

C11-2-07A Experiment to develop the relationship between the pressure and temperature of a gas using visual and graphical representations.
Include: historical contributions of Joseph Louis Gay-Lussac

Lesson 9: Gas Pressure and Temperature, Part B

C11-2-07B Experiment to develop the relationship between the pressure and temperature of a gas using numeric representation.
Include: historical contributions of Joseph Louis Gay-Lussac

Lesson 10: The Combined Gas Law

C11-2-08A Solve quantitative problems involving the relationships among the pressure, temperature, and volume of a gas using dimensional analysis.
Include: symbolic relationships

Lesson 11: Applications of the Gas Laws

C11-2-09 Identify various industrial, environmental, and recreational applications of gases.
Examples: self-contained underwater breathing apparatus (scuba), anaesthetics, air bags, acetylene welding, propane appliances, hyperbaric chambers…

Module 3: Chemical Reactions

Lesson 1: Atomic Masses and Isotopes

C11-3-01 Determine average atomic mass using isotopes and their relative abundance.
Include: atomic mass unit (amu)

C11-3-02 Research the importance and applications of isotopes.
Examples: nuclear medicine, stable isotopes in climatology, dating techniques…

Lesson 2: Polyatomic Compounds

C11-3-03 Write formulas and names for polyatomic compounds using International Union of Pure and Applied Chemistry (IUPAC) nomenclature.

Lesson 3: Atomic Mass Units

C11-3-04 Calculate the mass of compounds in atomic mass units.

Lesson 4: Balancing Chemical Equations

C11-3-05 Write and classify balanced chemical equations from written descriptions of reactions.
Include: polyatomic ions

Lesson 5: Chemical Reactions

C11-3-06 Predict the products of chemical reactions, given the reactants and type of reaction.
Include: polyatomic ions

Lesson 6: The Mole

C11-3-07 Describe the concept of the mole and its importance to measurement in chemistry.

C11-3-08 Calculate the molar mass of various substances.

Lesson 7: Molar Volume

C11-3-09 Calculate the volume of a given mass of a gaseous substance from its density at a given temperature and pressure.
Include: molar volume calculation

Lesson 8: Interconversions

C11-3-10 Solve problems requiring interconversions between moles, mass, volume, and number of particles.

Lesson 9: Empirical and Molecular Formulas

C11-3-11 Determine empirical and molecular formulas from percent composition or mass data.

Module 4: Stoichiometry

Lesson 1: Interpreting a Balanced Equation

C11-3-12 Interpret a balanced equation in terms of moles, mass, and volumes of gases.

Lesson 2: Stoichiometry, Part 1

C11-3-13 Solve stoichiometric problems involving moles, mass, and volume, given the reactants and products in a balanced equation.
Include: heat of reaction problems

Lesson 3: Stoichiometry, Part 2

C11-3-13 Solve stoichiometric problems involving moles, mass, and volume, given the reactants and products in a balanced equation.
Include: heat of reaction problems

Lesson 4: Limiting Reactant

C11-3-14 Identify the limiting reactant and calculate the mass of a product, given the reaction equation and reactant data.

Lesson 5: Laboratory Experiment-Limiting Reactant Investigation

C11-3-15 Perform a lab involving mass-mass or mass-volume relations, identifying the limiting reactant and calculating the mole ratio.
Include: theoretical yield, experimental yield

Lesson 6: Stoichiometry Applications

C11-3-16 Discuss the importance of stoichiometry in industry and describe specific applications.
Examples: analytical chemistry, chemical engineering, industrial chemistry…

Module 5: Solutions

Lesson 1: Types of Solutions

C11-4-01 Describe and give examples of various types of solutions.
Include: all nine possible types

Lesson 2: The Structure of Water

C11-4-02 Describe the structure of water in terms of electronegativity and the polarity of its chemical bonds.

Lesson 3: The Solution Process

C11-4-03 Explain the solution process of simple ionic and covalent compounds, using visual, particulate representations and chemical equations.
Include: crystal structure, dissociation, hydration

C11-4-04 Explain heat of solution with reference to specific applications.
Examples: cold packs, hot packs….

Lesson 4: Dissolving Polar and Non-Polar Substances

C11-4-05 Perform a lab to illustrate the formation of solutions in terms of the polar and non-polar nature of substances.
Include: soluble, insoluble, miscible, immiscible

Lesson 5: The Solubility Curve

C11-4-06 Construct, from experimental data, a solubility curve of a pure substance in water.

C11-4-07 Differentiate among saturated, unsaturated, and supersaturated solutions.

Lesson 6: Solving Solubility Problems

C11-4-08 Use a graph of solubility data to solve problems.

Lesson 7: Solubility of Gases

C11-4-09 Explain how a change in temperature affects the solubility of gases.

C11-4-10 Explain how a change in pressure affects the solubility of gases.

Lesson 8: Freezing Point Depression and Boiling Point Elevation

C11-4-11 Perform a lab to demonstrate freezing-point depression and boiling-point elevation.

C11-4-12 Explain freezing-point depression and boiling-point elevation at the molecular level.
Examples: antifreeze, road salt…

Lesson 9: Concentration

C11-4-13 Differentiate among, and give examples of, the use of various representations of concentration.

C11-4-14 Solve problems involving calculation for concentration, moles, mass, and volume.

Lesson 10: Preparing a Solution

C11-4-15 Prepare a solution, given the amount of solute (in grams) and the volume of solution (in millilitres), and determine the concentration in moles/litre.

Lesson 11: Dilutions

C11-4-16 Solve problems involving the dilution of solutions.
Include: dilution of stock solutions, mixing common solutions with different volumes and concentrations

C11-4-17 Perform a dilution from a solution of known concentration.

Lesson 12: Concentration Applications

C11-4-18 Describe examples of situations where solutions of known concentration are important.
Examples: pharmaceutical preparations, administration of drugs, aquariums, swimming pool disinfectants, gas mixes for scuba, radiator antifreeze…

Lesson 13: Our Water Supply

C11-4-19 Describe the process of treating a water supply, identifying the allowable concentration of metallic and organic species in water suitable for consumption.

Module 6: Organic Chemistry

Lesson 1: The Chemistry of Carbon

C11-5-01 Compare and contrast inorganic and organic chemistry.
Include: the contributions of Friedrich Wöhler and the overturn of vitalism

C11-5-02 Identify the origins and major sources of hydrocarbons and other organic compounds.
Include: natural and synthetic sources

Lesson 2: Hydrocarbons

C11-5-03 Describe the structural characteristics of carbon.
Include: bonding characteristics of the carbon atom in hydrocarbons (single, double, triple bonds)

C11-5-04 Compare and contrast the molecular structures of alkanes, alkenes, and alkynes.
Include: trends in melting points and boiling points of alkanes only

Lesson 3: Alkanes

C11-5-05 Name, draw, and construct structural models of the first 10 alkanes.
Include: IUPAC nomenclature, structural formulas, condensed structural formulas, molecular formulas, general formula $C_nH_{(2n+2)}$

C11-5-06 Name, draw, and construct structural models of the branched alkanes.
Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature

Lesson 4: Isomers of Alkanes

C11-5-07 Name, draw, and construct structural models of isomers for alkanes up to six-carbon atoms.
Include: condensed structural formulas

Lesson 5: Alkenes

C11-5-08 Outline the transformation of alkanes to alkenes and vice-versa.
Include: dehydrogenation/hydrogenation, molecular models

C11-5-09 Name, draw, and construct molecular models of alkenes and branched alkenes.
Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature, structural formulas, condensed structural formulas, molecular formulas, general formula C_nH_{2n}

C11-5-10 Differentiate between saturated and unsaturated hydrocarbons.

Lesson 6: Alkynes

C11-5-11 Outline the transformation of alkenes to alkynes and vice-versa.
Include: dehydrogenation/hydrogenation, molecular models

C11-5-12 Name, draw, and construct molecular models of alkynes and branched alkynes.
Include: six-carbon parent chain, methyl and ethyl substituent groups, IUPAC nomenclature, structural formulas, condensed structural formulas, molecular formulas, general formula C_nH_{2n-2}

Lesson 7: Aromatic Hydrocarbons

C11-5-13 Compare and contrast the aromatic and aliphatic hydrocarbons.
Include: molecular models, condensed structural formulas

C11-5-14 Describe uses of aromatic hydrocarbons.
Examples: polychlorinated biphenyls, caffeine, steroids, organic solvents (toluene, xylene)…

Lesson 8: Common Alcohols

C11-5-15 Write condensed structural formulas for and name common alcohols.
Include: maximum of six-carbon parent chain, IUPAC nomenclature

C11-5-16 Describe uses of methyl, ethyl, and isopropyl alcohols

Lesson 9: Organic Acids

C11-5-17 Write condensed structural formulas for and name organic acids.
Include: maximum of six-carbon parent chain, IUPAC nomenclature

C11-5-18 Describe uses or functions of common organic acids.
Examples: acetic, ascorbic, citric, formic, acetylsalicylic (ASA), lactic…

Lesson 10: Esters

C11-5-20 Write condensed structural formulas for and name esters.
Include: up to 6-C alcohols and 6-C organic acids, IUPAC nomenclature

C11-5-21 Describe uses of common esters.
Examples: pheromones, artificial flavourings…

Lesson 11: Chemistry and Industry

C11-5-22 Describe the process of polymerization and identify important natural and synthetic polymers.
Examples: polyethylene, polypropylene, polystyrene, polytetrafluoroethylene (Teflon®)

C11-5-23 Describe how the products of organic chemistry have influenced quality of life.
Examples: synthetic rubber, nylon, medicines, polytetrafluoroethylene (Teflon®)

Lesson 12: Issues Related to Organic Chemistry

C11-5-24 Use the decision-making process to investigate an issue related to organic chemistry.
Examples: gasohol production, alternative energy sources, recycling of plastics…

NOTES

Grade 11 Chemistry (30S)

Final Practice Examination

GRADE 11 CHEMISTRY (30S)

Final Practice Examination

IInstructions

The final examination will be weighted as follows

Modules 1–3	15–20%
Modules 4–6	80–85%

The format of the examination will be as follows:

Part A: Fill-in-the-Blanks	22 x 1 = 22 marks
Part B: Multiple Choice	46 x 1 = 46 marks
Part C: Short Answer	32 marks
Total Marks	100 marks

Include units with all answers as required.

Useful Information

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- some scrap paper
- a ruler
- a scientific or graphing calculator

You will have a maximum of 2.5 hours to complete your final exam.

Part A: Fill-in-the-Blanks (22 Marks)

Use the Word Bank at the end of this exam to help you complete the "Fill in the Blank" questions. As each blank is worth one mark, some questions will have a total value of two marks. Note that there are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use. The same term may be used more than once in this section.

Stoichiometry ***(4 marks)***

1. The calculated amount of product formed during a reaction is called the ____________________ yield.
2. For a given chemical reaction, the actual yield is always ____________________ than the theoretical yield.
3. For ____________________ changes, the release of energy is represented by writing the energy term as a product when writing a chemical equation.
4. Balanced chemical equations contain important information about the amount of reactants required to produce given products. These amounts are represented by ____________________.

Solutions ***(10 marks)***

5. A ____________________ is defined as a mixture of two or more substances that are evenly distributed.
6. The shape of the water molecule, combined with the nature of its bonds, makes water a ____________________ molecule.
7. The total heat change in the dissolving process is called the heat of ____________________.
8. Non-polar substances, like waxes and oils, are ____________________ in water.
9. If a solution could dissolve more solute at a particular temperature, the solution is ____________________.
10. The number of moles of solute dissolved in 1 L of solution is known as ____________________.
11. The attraction an atom has for the shared electrons in a covalent bond is called ____________________ .
12. When water surrounds individual molecules or ions, the molecules or ions are said to be ____________________.

13. Polar and charged substances dissolve well in ____________________ solvents because of the electrostatic attraction between opposite charges.

14. Adding a solute to a solvent lowers the ____________________ pressure of the solvent.

Organic Chemistry ***(8 marks)***

15. Inorganic compounds do NOT tend to contain ____________________.

16. Decaying animals and vegetation is a major source of ____________________ compounds.

17. When carbon atoms are bonded together in a tetrahedral lattice arrangement, ____________________ is created.

18. Any alkane with one or more alkyl groups is automatically a ____________________ alkane.

19. Ethene, a simple alkene, can be transformed back into ethane by adding ____________________.

20. All hydrocarbons that do not possess rings are called ____________________ compounds.

21. The functional group that identifies a compound as an alcohol is called the ____________________ group.

22. The process of forming an ester from a reaction between an organic acid and an ____________________ is called esterification.

Part B: Multiple Choice (46 Marks)

For each Multiple Choice question, shade in the circle that corresponds to your answer on the Bubble Sheet at the end of this exam. DO NOT circle your answers directly on the exam.

Stoichiometry ***(6 marks)***

1. Which of these interpretations of the following balanced equation is TRUE?

   $$2S_{(s)} + 3O_{2(g)} \rightarrow 2SO_{3(g)}$$

   a) 2 atoms of S and 3 atoms of O_2 form 2 atoms of SO_3
   b) 2 grams of S and 3 grams of O_2 form 2 grams of SO_3
   c) 2 moles of S and 3 moles of O_2 form 2 moles of SO_3
   d) 2 L of S and 3 L of O_2 form 2 L of SO_3

2. Which type of stoichiometric calculation does not involve the gram formula mass?
   a) Mass-mass problems
   b) Mass-particle problems
   c) Mass-volume problems
   d) Volume-volume problems

3. The ratio of the actual yield to the theoretical yield is known as the
   a) Excess yield
   b) Reagent yield
   c) Percent yield
   d) Experimental yield

4. In the following balanced equation, how many moles of aluminum are needed to form 3.70 moles of aluminum oxide, Al_2O_3?

   $$4Al_{(s)} + O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

   a) 7.40 moles
   b) 3.70 moles
   c) 2.00 moles
   d) 1.85 moles

5. Convert 35.0 L of nitrogen gas to moles of nitrogen gas at STP.
   a) 1.56 moles
   b) 0.640 moles
   c) 7.84 moles
   d) 22.4 moles

6. Which of the following quantities is conserved in *every* chemical reaction?
   a) Molecules
   b) Mass
   c) Formula units
   d) Moles

Solutions *(20 marks)*

7. Which of the following is LESS soluble in hot water than in cold water?
   a) CO_2
   b) NaCl
   c) $NaNO_3$
   d) KBr

8. What can be done to crystallize a supersaturated solution?
   a) Heat the solution.
   b) It will crystallize if you leave it alone.
   c) Add a crystal of the solute or scratch the glass.
   d) Expose the solution to ultraviolet light.

9. In a concentrated solution, there is
   a) No solvent.
   b) A large amount of solute.
   c) A small amount of solvent.
   d) No solute.

10. In which of the following is concentration expressed in percent by volume?
    a) 10% (v/v)
    b) 10% (m/v)
    c) 10% (m/m)
    d) 10%

11. Which of the following is NOT a colligative property of a solution?
    a) Boiling point elevation
    b) Freezing point depression
    c) Vapour pressure lowering
    d) Solution saturation

12. What is the maximum amount of KCl that can be dissolved into 150.0 g of water? (The solubility of KCl is 34.0 g/100 mL at STP.)
    a) 51.0 g
    b) 22.7 g
    c) 34.0 g
    d) 5.10 g

13. Which of the following pairs of substances are miscible?
    a) Water and gasoline
    b) Water and salt (NaCl)
    c) Water and oxygen
    d) Water and ethanol (alcohol)

14. At STP, the solubility of solute XY is $\dfrac{10\text{ g}}{100\text{ g water}}$. Which of the following solution concentrations could represent an *unsaturated* solution of solute XY?

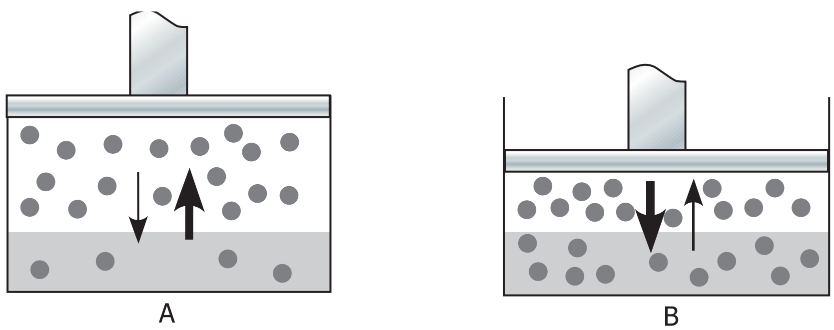
a) $\dfrac{10\text{ g}}{100\text{ g water}}$

b) $\dfrac{9\text{ g}}{100\text{ g water}}$

c) $\dfrac{5\text{ g}}{50\text{ g water}}$

d) $\dfrac{11\text{ g}}{100\text{ g water}}$

15. Use the following two diagrams of a gas-liquid solution to help you determine which statement below is FALSE.



a) The increased pressure in diagram B illustrates an increased solubility of the gas in the liquid.
b) The increased pressure shown in diagram B forces the gas into contact with the liquid.
c) Diagram A shows a greater amount of gas in solution, whereby the liquid holds onto the gas particles.
d) When the pressure is reduced in diagram A, the solubility of the dissolved gas is reduced.

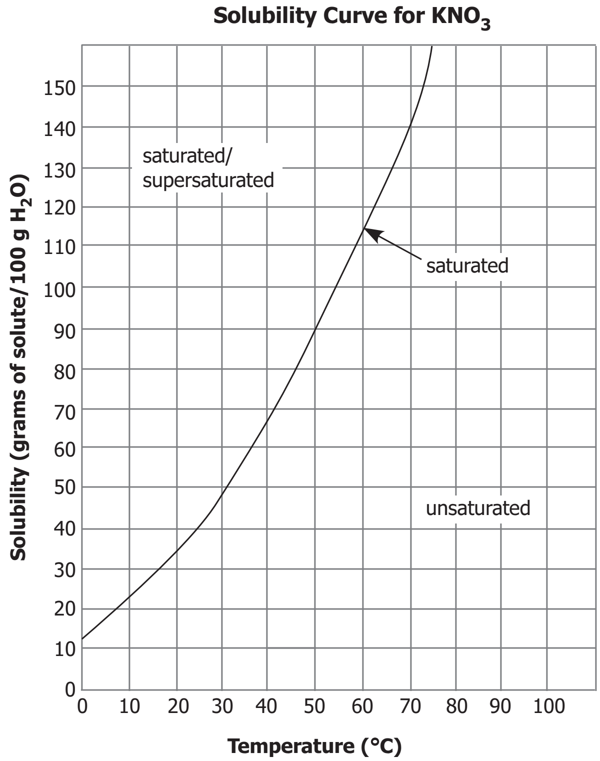
16. The solubility of a gas in a liquid
a) Increases as the pressure of the gas above the liquid increases.
b) Decreases as the pressure of the gas above the liquid increases.
b) Increases as the pressure of the gas above the liquid decreases.
d) Is unrelated to the pressure of the gas above the liquid.

17. Which type of mixture could most likely be filtered using filter paper?
    a) A colloid
    b) A suspension
    c) A solution
    d) An emulsion

18. Which of these statements regarding the water molecule is FALSE?
    a) Oxygen is more electronegative than the hydrogen.
    b) The electrons between the hydrogen and oxygen atoms in each bond lie more towards the oxygen than they do towards the hydrogen.
    c) The hydrogen atoms are bonded to the oxygen at an angle of 104.5°, which gives the water molecule its characteristic bent shape.
    d) The water molecule is a non-polar molecule.

19. Which statement below would NOT ensure greater conductivity of an electric current?
    a) There must be charged particles or ions present in the solution.
    b) Particles must move freely through the solution.
    c) There must be fewer ions present in solution.
    d) There must be a lower volume of solvent in which the ions are dissolved.

Use the "Solubility Curve for KNO_3" below for Questions 20 to 22.



20. Estimate the approximate solubility of KNO_3 at 30ºC.
   a) 16 g/100 g H_2O
   b) 33 g/100 g H_2O
   c) 48 g/100 g H_2O
   d) 60 g/100 g H_2O

21. Estimate the temperature at which the solubility of potassium nitrate is 50 g/100 g.
   a) About 90ºC
   b) About 20ºC
   c) About 30ºC
   d) About 8ºC

22. Indicate which of the following sets of data represents a *saturated* solution of potassium nitrate.
   a) 25ºC: 40 g/100 g H_2O
   b) 63ºC: 140 g/100 g H_2O
   c) 8ºC: 10 g/100 g H_2O
   d) 70ºC: 150 g/100 g H_2O

23. How many moles of NaOH would be needed to make 0.0500 L of a 0.750 mol/L solution?
    a) 15.0 mol
    b) 0.0375 mol
    c) 50.0 mol
    d) 0.750 mol

24. Identify the FINAL step to follow when preparing a solution.
    a) Mass out the solute and add it to the flask.
    b) Add more solvent until you reach the required amount.
    c) Mass out the solvent and add it to the flask.
    d) Add about half the required volume of solvent to the flask.

25. You start with a solution that is 0.800 mol/L and exactly 0.0700 L. You need to prepare a 0.300 mol/L solution. What is the final volume of the solution?
    a) 3.43 L
    b) 0.026 L
    c) 0.580 L
    d) 0.187 L

26. Which method of water treatment is useful for controlling disease-causing organisms such as viruses, bacteria, and parasites?
    a) Water softening
    b) Filtration
    c) Chlorination
    d) Distillation

Organic Chemistry (20 marks)

27. Which of these statements does NOT accurately describe tar sands?
    a) Tar sands provide a synthetically produced source of hydrocarbons.
    b) Tar sands are a combination of clay, sand, water, and bitumen.
    c) Tar sands can be mined and processed to extract the oil-rich bitumen.
    d) Bitumen requires no further refining and can be pumped from the ground in its natural state.

28. Which type of bond will carbon commonly form?
    a) Covalent
    b) Ionic
    c) Metallic
    d) None of these

29. All of these hydrocarbons are *unsaturated* except for
    a) Benzene
    b) Alkenes
    c) Alkanes
    d) Alkynes

30. Which of the following is the correct condensed structural formula for *butane*?
    a) $CH_3(CH_2)_3CH_3$
    b) $CH_3(CH_2)_2CH_3$
    c) $(CH_3)_3CH_3$
    d) C_4H_{10}

31. Name the following alkane: C_7H_{16}
    a) Heptane
    b) Hexane
    c) Decane
    d) Octane

32. The correct structural formula for *2,2–dimethylhexane* is:

a)
$$\begin{array}{l} \qquad\quad\; CH_2CH_3 \\ \qquad\quad\;\; | \\ CH_3-C-CH_2-CH_2-CH_2-CH_3 \\ \qquad\quad\;\; | \\ \qquad\quad\; CH_3 \end{array}$$

b)
$$\begin{array}{l} \qquad\quad\; CH_3 \\ \qquad\quad\;\; | \\ CH_3-CH-CH_2-CH_2-CH_2-CH_2-CH_3 \\ \qquad\quad\;\; | \\ \qquad\quad\; CH_3 \end{array}$$

c)
$$\begin{array}{l} CH_3 \\ \;| \\ CH-CH_2-CH_2-CH_2-CH_2-CH_3 \\ \;| \\ CH_3 \end{array}$$

d)
$$\begin{array}{l} \qquad\quad\; CH_3 \\ \qquad\quad\;\; | \\ CH_3-C-CH_2-CH_2-CH_2-CH_3 \\ \qquad\quad\;\; | \\ \qquad\quad\; CH_3 \end{array}$$

33. Molecules that have the same molecular formula but different structural formulas are called

a) Allotropes

b) Sterioisomers

c) Structural isomers

d) Isotopes

34. Which straight-chain alkane is a structural isomer of *3-propylheptane*?

a) 10 carbon atoms = decane

b) 6 carbon atoms = hexane

c) 9 carbon atoms = nonane

d) 5 carbon atoms = pentane

35. The correct name for the alkene $CH_3CH{=}CH_2$ is
  a) propene
  b) prop–2–ene
  c) 2–propene
  d) prop–3–ene

36. The correct structural formula for *2–methylbut–1–ene* is
  a)
$$\begin{array}{ccccccc} & & & & CH_3 & & \\ & & & & | & & \\ CH_3 & - & CH & - & CH & = & CH_3 \end{array}$$
  b)
$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_2 & = & C & - & CH_2 & - & CH_3 \end{array}$$
  c)
$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_3 & - & C & = & CH & - & CH_3 \end{array}$$
  d)
$$\begin{array}{ccccccc} & & & & & & CH_3 \\ & & & & & & | \\ CH_3 & - & CH & = & CH_2 & - & CH_2 \end{array}$$

37. The correct name for the alkyne $CH{\equiv}CCH_2CH_2CH_3$ is
  a) pent–1–yne
  b) 1–pentyne
  c) pentyne
  d) pent–4–yne

38. The correct structural formula for *3–ethylpent–1–yne* is

a)
$$\begin{array}{ccccccccc} & & & & \mathrm{CH_2CH_3} & & & & \\ & & & & | & & & & \\ \mathrm{C} & \equiv & \mathrm{C} & - & \mathrm{C} & - & \mathrm{CH_2} & - & \mathrm{CH_3}\end{array}$$

b)
$$\begin{array}{ccccccccc} & & \mathrm{CH_3} & & & & & & \\ & & | & & & & & & \\ \mathrm{CH} & \equiv & \mathrm{C} & - & \mathrm{CH} & - & \mathrm{CH_2} & - & \mathrm{CH_3}\end{array}$$

c)
$$\begin{array}{ccccccccc} \mathrm{CH_2CH_3} & & & & & & & & \\ | & & & & & & & & \\ \mathrm{CH_2} & - & \mathrm{CH_2} & - & \mathrm{C} & \equiv & \mathrm{C} & - & \mathrm{CH_3}\end{array}$$

d)
$$\begin{array}{ccccccccc} & & & & \mathrm{CH_2CH_3} & & & & \\ & & & & | & & & & \\ \mathrm{CH} & \equiv & \mathrm{C} & - & \mathrm{CH} & - & \mathrm{CH_2} & - & \mathrm{CH_3}\end{array}$$

39. Give the IUPAC name for the alcohol $CH_3CH(OH)CH_2CH(CH_3)CH_3$.
   a) 2–methylpentan–4–ol
   b) 2–methyl–4–pentanol
   c) 4–methylpentan–2–ol
   d) 4–methyl–2–pentanol

40. Identify which of the following alcohols is most likely used to make hand soap.
   a) Isopropyl
   b) Glycerol
   c) Ethanol
   d) Methanol

41. Name the following carboxylic acid: $CH_3CH_2CH(CH_3)CH_2COOH$
   a) 3–methylpentanoic acid
   b) 2–methylbutanoic acid
   c) methyl–3–pentanoic acid
   d) methyl–2–butanoic acid

42 Identify which of the following carboxylic acids is responsible for the sting in ant bites.

a) Benzoic acid
b) Lactic acid
c) Formic acid
d) Acetic acid

43. Name the ester $CH_3CH_2CH_2CH_2COOCH_3$.

a) methyl butanoate
b) methylpentanoic acid
c) pentyl methanoate
d) methyl pentanoate

44. A very large molecule made of many smaller repeating units is known as

a) A monomer
b) A polymer
c) An ester
d) An allotrope

45. Which of these examples of is NOT a polymer formed by an addition reaction?

a) Graphite
b) Teflon™
c) Polypropylene
d) Synthetic rubber

46. Which of these polymers is used for moulded plastics and film?

a) Polyethylene
b) Polyvinyl chloride (PVC)
c) Teflon™
d) Polypropylene

Part C: Short Answer (32 Marks)

Answer each of the questions below using the space provided. Pay attention to the number of marks that each question is worth, as this may help you decide how much information to provide for full marks. For questions that involve calculations, show your work and check your final answer for the correct number of significant figures and the appropriate unit.

Stoichiometry ***(15 marks)***

1. How many moles of $O_{2(g)}$ react with 2.4 moles of Fe in the following rusting reaction? *(2 marks)*

$$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$$

2. What quantity of heat is produced in the complete combustion of 60.2 g of ethane gas (C_2H_6), according to the following balanced chemical reaction? The heat of combustion of ethane is 1560 kJ/mol and its molar mass is 30.0 g/mol. *(4 marks)*

$$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(g)}$$

3. How many grams of CO_2 would be produced if 45 g of $C_6H_{12}O_6$ (glucose) reacted completely with oxygen? Glucose = 180.0 g/mol; CO_2 = 44.0 g/mol. *(4 marks)*

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$$

4. Given 5.0 moles of sulfur and 8.4 moles of oxygen gas, as well as

$$2S + 3O_2 \rightarrow 2SO_3$$

a) Identify the limiting factor and the excess reactant. *(3 marks)*

b) Calculate the moles of excess reactant that remain. *(2 marks)*

Solutions ***(7 marks)***

5. Describe any three properties that are true of a solution in terms of the particulate view of matter. *(3 marks)*

6. Write the equation for dissolving $Ag_2CrO_{4(s)}$ in water. *(2 marks)*

7. What is the number of moles of solute in 0.650 L of a 0.40 mol/L solution? *(2 marks)*

Organic Chemistry ***(10 marks)***

8. Draw the structural formulas for the following hydrocarbons. *(2 marks x 3 = 6 marks)*

   a) 2,4–dimethylpentane

   b) hex–2–ene

   c) 4–ethylhex–2–yne

9. Complete the following table. *(4 marks)*

	Aliphatic Hydrocarbon	**Aromatic Hydrocarbon**
Similarities		
Differences		

NOTES

Grade 11 Chemistry
Final Practice Examination

Word Bank

Use the following word bank to help you complete the "Fill-in-the-Blank" portion of your Final Examination. Note that there may be MORE terms here than you need, so read over the list carefully before choosing the terms that you want to use. You can also use certain words more than once.

alcohol(s)
aliphatic
alkyl
allotrope
amorphous carbon
aqueous
aromatics
benzene
boiling
branched-chain
buckminsterfullerene
carbon
carboxyl
carboxylic
coefficients
colligative
concentration
cracking
crude
decrease(s)
dehydrogenation
depression
diamond
dilution
electronegativity
electrons
electrostatic
emulsion
endothermic
equal
esterification
excess
exothermic
graphite
greater
higher
hydrogenation
hydrated
hydrocarbon
hydrogen
hydroxyl
immiscible
increase(s)
inorganic
insoluble
isomers
less
limiting
miscible
molarity
moles
more
non-polar
number
organic
percent
phenyl
polar
products
proportion(s)
ratio(s)
R-COOH
reactant
saturated
smaller
soluble
solute
solution
solvation
solvent
stoichiometry
substituent
supersaturated
suspensions
theoretical
unsaturated
vapour
water

Grade 11 Chemistry
Final Practice Examination
Bubble Sheet

Name: ______________________________________ ______________ / 46

For each Multiple Choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the exam.

	A	B	C	D		A	B	C	D		A	B	C	D		A	B	C	D
1.	❍	❍	❍	❍	14.	❍	❍	❍	❍	27.	❍	❍	❍	❍	40.	❍	❍	❍	❍
2.	❍	❍	❍	❍	15.	❍	❍	❍	❍	28.	❍	❍	❍	❍	41.	❍	❍	❍	❍
3.	❍	❍	❍	❍	16.	❍	❍	❍	❍	29.	❍	❍	❍	❍	42.	❍	❍	❍	❍
4.	❍	❍	❍	❍	17.	❍	❍	❍	❍	30.	❍	❍	❍	❍	43.	❍	❍	❍	❍
5.	❍	❍	❍	❍	18.	❍	❍	❍	❍	31.	❍	❍	❍	❍	44.	❍	❍	❍	❍
6.	❍	❍	❍	❍	19.	❍	❍	❍	❍	32.	❍	❍	❍	❍	45.	❍	❍	❍	❍
7.	❍	❍	❍	❍	20.	❍	❍	❍	❍	33.	❍	❍	❍	❍	46.	❍	❍	❍	❍
8.	❍	❍	❍	❍	21.	❍	❍	❍	❍	34.	❍	❍	❍	❍					
9.	❍	❍	❍	❍	22.	❍	❍	❍	❍	35.	❍	❍	❍	❍					
10.	❍	❍	❍	❍	23.	❍	❍	❍	❍	36.	❍	❍	❍	❍					
11.	❍	❍	❍	❍	24.	❍	❍	❍	❍	37.	❍	❍	❍	❍					
12.	❍	❍	❍	❍	25.	❍	❍	❍	❍	38.	❍	❍	❍	❍					
13.	❍	❍	❍	❍	26.	❍	❍	❍	❍	39.	❍	❍	❍	❍					

Periodic Table of the Elements

Atomic Number	19
Symbol	**K**
Name	Potassium
Relative Atomic Mass	39.1

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Ununhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Electronegativities Table

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 **Lanthanide Series**	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 **Actinide Series**	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Alphabetical Listing of the Elements and Their Atomic Masses

Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Actinium	(227)	Gold	197.0	Praseodymium	140.9
Aluminum	27.0	Hafnium	178.5	Promethium	(145)
Americium	(243)	Hassium	(265)	Protactinum	(231)
Antimony	121.7	Helium	4.0	Radium	(226)
Argon	39.9	Holmium	164.9	Radon	(222)
Arsenic	74.9	Hydrogen	1.0	Rhenium	186.2
Astatine	(210)	Indium	114.8	Rhodium	102.9
Barium	137.3	Iodine	126.9	Rubidium	85.5
Berkelium	(247)	Iridium	192.2	Ruthenium	101.1
Beryllium	9.0	Iron	55.8	Rutherfordium	(261)
Bismuth	209.0	Krypton	83.8	Samarium	150.4
Bohrium	(264)	Lanthanum	138.9	Scandium	45.0
Boron	10.8	Lawrencium	(257)	Seaborgium	(263)
Bromine	79.9	Lead	207.2	Selenium	79.0
Cadmium	112.4	Lithium	6.9	Silicon	28.1
Calcium	40.1	Lutetium	175.0	Silver	107.9
Californium	(251)	Magnesium	24.3	Sodium	23.0
Carbon	12.0	Manganese	54.9	Strontium	87.6
Cerium	140.1	Meitnerium	(266)	Sulfur	32.1
Cesium	132.9	Mendelevium	(256)	Tantalum	180.9
Chlorine	35.5	Mercury	200.6	Technetium	(98)
Chromium	52.0	Molybdenum	95.9	Tellurium	127.6
Cobalt	58.9	Neodymium	144.2	Terbium	158.9
Copernicium	(277)	Neon	20.2	Thallium	204.4
Copper	63.5	Neptunium	(237)	Thorium	232.0
Curium	(247)	Nickel	58.7	Thulium	168.9
Dubnium	(262)	Niobium	92.9	Tin	118.7
Dysprosium	162.5	Nitrogen	14.0	Titanium	47.9
Einsteinium	(254)	Nobelium	(259)	Tungsten	183.8
Erbium	167.3	Osmium	190.2	Uranium	238.0
Europium	152.0	Oxygen	16.0	Vanadium	50.9
Fermium	(257)	Palladium	106.4	Xenon	131.3
Fluorine	19.0	Phosphorus	31.0	Ytterbium	173.0
Francium	(223)	Platinum	195.1	Yttrium	88.9
Gadolinium	157.2	Plutonium	(244)	Zinc	65.4
Gallium	69.7	Polonium	(209)	Zirconium	91.2
Germanium	72.6	Potassium	39.1		

Names, Formulas, and Charges of Common Ions

Positive Ions (Cations)

Name	Symbol	Name	Symbol
aluminum	Al^{3+}	magnesium	Mg^{2+}
ammonium	NH_4^+	**manganese(II)**	$\mathbf{Mn^{2+}}$
barium	Ba^{2+}	manganese(IV)	Mn^{4+}
cadmium	Cd^{2+}	mercury(I)	Hg_2^{2+}
calcium	Ca^{2+}	mercury(II)	Hg^{2+}
chromium(II)	Cr^{2+}	**nickel(II)**	$\mathbf{Ni^{2+}}$
chromium(III)	$\mathbf{Cr^{3+}}$	nickel(III)	Ni^{3+}
copper(I)	Cu^+	potassium	K^+
copper(II)	$\mathbf{Cu^{2+}}$	silver	Ag^+
hydrogen	H^+	sodium	Na^+
iron(II)	Fe^{2+}	strontium	Sr^{2+}
iron(III)	$\mathbf{Fe^{3+}}$	tin(II)	Sn^{2+}
lead(II)	$\mathbf{Pb^{2+}}$	**tin(IV)**	$\mathbf{Sn^{4+}}$
lead(IV)	Pb^{4+}	zinc	Zn^{2+}
lithium	Li^+		

continued

Negative Ions (Anions)

Name	Symbol	Name	Symbol
acetate	$C_2H_3O_2^-$ (CH_3COO^-)	nitrate	NO_3^-
azide	N_3^-	nitride	N^{3-}
bromide	Br^-	nitrite	NO_2^-
bromate	BrO_3^-	oxalate	$C_2O_4^{2-}$
carbonate	CO_3^{2-}	hydrogen oxalate	$HC_2O_4^-$
hydride	H^-	oxide	O^{2-}
hydrogen carbonate or bicarbonate	HCO_3^-	perchlorate	ClO_4^-
chlorate	ClO_3^-	permanganate	MnO_4^-
chloride	Cl^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}
chromate	CrO_4^{2-}	dihydrogen phosphate	$H_2PO_4^-$
citrate	$C_6H_5O_7^{3-}$	silicate	SiO_3^{2-}
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfate	HSO_4^-
fluoride	F^-	sulfide	S^{2-}
hydroxide	OH^-	hydrogen sulfide	HS^-
hypochlorite	ClO^-	sulfite	SO_3^{2-}
iodide	I^-	hydrogen sulfite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4^+ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

continued

Anions (Negative Ions)

1^- charge		**1^- charge**		**2^- charge**	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^-	Bromate	OH^-	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^-	Bromide	IO_3^-	Iodate	O^{2-}	Oxide
ClO_3^-	Chlorate	I^-	Iodide	O_2^{2-}	Peroxide
Cl^-	Chloride	NO_3^-	Nitrate	SO_4^{2-}	Sulfate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulfide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulfite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocyanate	**3^- charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Grade 11 Chemistry (30S)

Final Practice Examination

Answer Key

Grade 11 Chemistry (30S)

Final Practice Examination
Answer Key

IInstructions

The final examination will be weighted as follows

Modules 1–3	15–20%
Modules 4–6	80–85%

The format of the examination will be as follows:

Part A: Fill-in-the-Blanks	22 x 1 = 22 marks
Part B: Multiple Choice	46 x 1 = 46 marks
Part C: Short Answer	32 marks
Total Marks	100 marks

Include units with all answers as required.

Useful Information

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- some scrap paper
- a ruler
- a scientific or graphing calculator

You will have a maximum of 2.5 hours to complete your final exam.

Part A: Fill-in-the-Blanks (22 Marks)

Use the Word Bank at the end of this exam to help you complete the "Fill in the Blank" questions. As each blank is worth one mark, some questions will have a total value of two marks. Note that there are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use. The same term may be used more than once in this section.

Stoichiometry ***(4 marks)***

1. The calculated amount of product formed during a reaction is called the ____________________ yield. ***Theoretical***
2. For a given chemical reaction, the actual yield is always ____________________ than the theoretical yield. ***Less / smaller***
3. For ____________________ changes, the release of energy is represented by writing the energy term as a product when writing a chemical equation. ***Exothermic***
4. Balanced chemical equations contain important information about the amount of reactants required to produce given products. These amounts are represented by ____________________. ***Coefficients***

Solutions ***(10 marks)***

5. A ____________________ is defined as a mixture of two or more substances that are evenly distributed. ***Solution***
6. The shape of the water molecule, combined with the nature of its bonds, makes water a ____________________ molecule. ***Polar***
7. The total heat change in the dissolving process is called the heat of ____________________. ***Solution***
8. Non-polar substances, like waxes and oils, are ____________________ in water. ***Insoluble / immiscible***
9. If a solution could dissolve more solute at a particular temperature, the solution is ____________________. ***Unsaturated***
10. The number of moles of solute dissolved in 1 L of solution is known as ____________________. ***Molarity / concentration***
11. The attraction an atom has for the shared electrons in a covalent bond is called ____________________ . ***Electronegativity***
12. When water surrounds individual molecules or ions, the molecules or ions are said to be ____________________. ***Hydrated***

13. Polar and charged substances dissolve well in ____________________ solvents because of the electrostatic attraction between opposite charges. ***Polar***

14. Adding a solute to a solvent lowers the ____________________ pressure of the solvent. ***Vapour***

Organic Chemistry ***(8 marks)***

15. Inorganic compounds do NOT tend to contain ____________________. ***Carbon***

16. Decaying animals and vegetation is a major source of ____________________ compounds. ***Hydrocarbon***

17. When carbon atoms are bonded together in a tetrahedral lattice arrangement, ____________________ is created. ***Diamond***

18. Any alkane with one or more alkyl groups is automatically a ____________________ alkane. ***Branched-chain***

19. Ethene, a simple alkene, can be transformed back into ethane by adding ____________________. ***Hydrogen***

20. All hydrocarbons that do not possess rings are called ____________________ compounds. ***Aliphatic***

21. The functional group that identifies a compound as an alcohol is called the ____________________ group. ***Hydroxyl***

22. The process of forming an ester from a reaction between an organic acid and an ____________________ is called esterification. ***Alcohol***

Part B: Multiple Choice (46 Marks)

For each Multiple Choice question, shade in the circle that corresponds to your answer on the Bubble Sheet at the end of this exam. DO NOT circle your answers directly on the exam.

Stoichiometry ***(6 marks)***

1. Which of these interpretations of the following balanced equation is TRUE?

   $$2S_{(s)} + 3O_{2(g)} \rightarrow 2SO_{3(g)}$$

   a) 2 atoms of S and 3 atoms of O_2 form 2 atoms of SO_3
   b) 2 grams of S and 3 grams of O_2 form 2 grams of SO_3
   c) 2 moles of S and 3 moles of O_2 form 2 moles of SO_3
   d) 2 L of S and 3 L of O_2 form 2 L of SO_3

2. Which type of stoichiometric calculation does not involve the gram formula mass?
   a) Mass-mass problems
   b) Mass-particle problems
   c) Mass-volume problems
   d) Volume-volume problems

3. The ratio of the actual yield to the theoretical yield is known as the
   a) Excess yield
   b) Reagent yield
   c) Percent yield
   d) Experimental yield

4. In the following balanced equation, how many moles of aluminum are needed to form 3.70 moles of aluminum oxide, Al_2O_3?

   $$4Al_{(s)} + O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

   a) 7.40 moles
   b) 3.70 moles
   c) 2.00 moles
   d) 1.85 moles

5. Convert 35.0 L of nitrogen gas to moles of nitrogen gas at STP.
   a) 1.56 moles
   b) 0.640 moles
   c) 7.84 moles
   d) 22.4 moles

6. Which of the following quantities is conserved in *every* chemical reaction?
   a) Molecules
   b) Mass
   c) Formula units
   d) Moles

Solutions *(20 marks)*

7. Which of the following is LESS soluble in hot water than in cold water?
   a) CO_2
   b) NaCl
   c) $NaNO_3$
   d) KBr

8. What can be done to crystallize a supersaturated solution?
   a) Heat the solution.
   b) It will crystallize if you leave it alone.
   c) Add a crystal of the solute or scratch the glass.
   d) Expose the solution to ultraviolet light.

9. In a concentrated solution, there is
   a) No solvent.
   b) A large amount of solute.
   c) A small amount of solvent.
   d) No solute.

10. In which of the following is concentration expressed in percent by volume?
    a) 10% (v/v)
    b) 10% (m/v)
    c) 10% (m/m)
    d) 10%

11. Which of the following is NOT a colligative property of a solution?
    a) Boiling point elevation
    b) Freezing point depression
    c) Vapour pressure lowering
    d) Solution saturation

12. What is the maximum amount of KCl that can be dissolved into 150.0 g of water? (The solubility of KCl is 34.0 g/100 mL at STP.)
    a) 51.0 g
    b) 22.7 g
    c) 34.0 g
    d) 5.10 g

13. Which of the following pairs of substances are miscible?
    a) Water and gasoline
    b) Water and salt (NaCl)
    c) Water and oxygen
    d) Water and ethanol (alcohol)

14. At STP, the solubility of solute XY is $\frac{10\text{ g}}{100\text{ g water}}$. Which of the following solution concentrations could represent an *unsaturated* solution of solute XY?

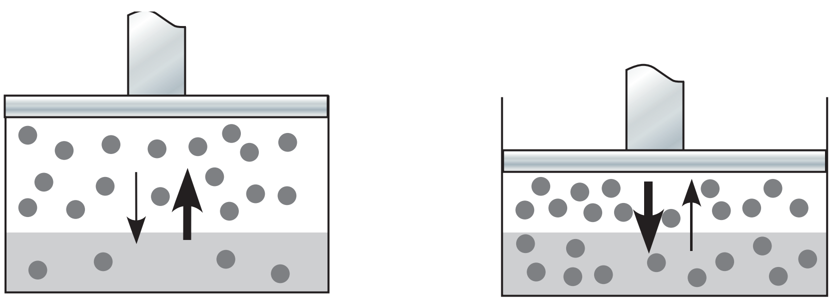
a) $\frac{10\text{ g}}{100\text{ g water}}$

b) $\frac{9\text{ g}}{100\text{ g water}}$

c) $\frac{5\text{ g}}{50\text{ g water}}$

d) $\frac{11\text{ g}}{100\text{ g water}}$

15. Use the following two diagrams of a gas-liquid solution to help you determine which statement below is FALSE.



a) The increased pressure in diagram B illustrates an increased solubility of the gas in the liquid.
b) The increased pressure shown in diagram B forces the gas into contact with the liquid.
c) Diagram A shows a greater amount of gas in solution, whereby the liquid holds onto the gas particles.
d) When the pressure is reduced in diagram A, the solubility of the dissolved gas is reduced.

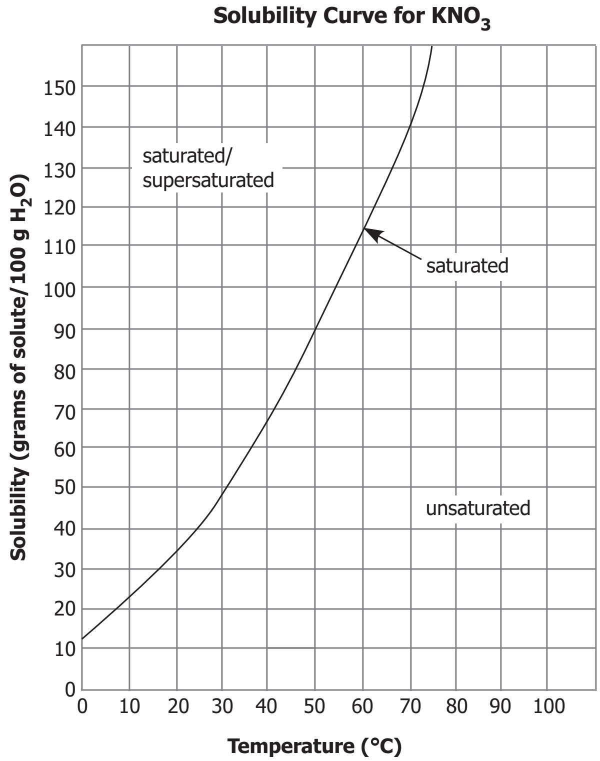
16. The solubility of a gas in a liquid
a) Increases as the pressure of the gas above the liquid increases.
b) Decreases as the pressure of the gas above the liquid increases.
b) Increases as the pressure of the gas above the liquid decreases.
d) Is unrelated to the pressure of the gas above the liquid.

17. Which type of mixture could most likely be filtered using filter paper?
    a) A colloid
    b) A suspension
    c) A solution
    d) An emulsion

18. Which of these statements regarding the water molecule is FALSE?
    a) Oxygen is more electronegative than the hydrogen.
    b) The electrons between the hydrogen and oxygen atoms in each bond lie more towards the oxygen than they do towards the hydrogen.
    c) The hydrogen atoms are bonded to the oxygen at an angle of 104.5°, which gives the water molecule its characteristic bent shape.
    d) The water molecule is a non-polar molecule.

19. Which statement below would NOT ensure greater conductivity of an electric current?
    a) There must be charged particles or ions present in the solution.
    b) Particles must move freely through the solution.
    c) There must be fewer ions present in solution.
    d) There must be a lower volume of solvent in which the ions are dissolved.

Use the "Solubility Curve for KNO_3" below for Questions 20 to 22.



20. Estimate the approximate solubility of KNO_3 at 30ºC.
   a) 16 g/100 g H_2O
   b) 33 g/100 g H_2O
   c) 48 g/100 g H_2O
   d) 60 g/100 g H_2O

21. Estimate the temperature at which the solubility of potassium nitrate is 50 g/100 g.
   a) About 90ºC
   b) About 20ºC
   c) About 30ºC
   d) About 8ºC

22. Indicate which of the following sets of data represents a *saturated* solution of potassium nitrate.
   a) 25ºC: 40 g/100 g H_2O
   b) 63ºC: 140 g/100 g H_2O
   c) 8ºC: 10 g/100 g H_2O
   d) 70ºC: 150 g/100 g H_2O

23. How many moles of NaOH would be needed to make 0.0500 L of a 0.750 mol/L solution?
    a) 15.0 mol
    b) 0.0375 mol
    c) 50.0 mol
    d) 0.750 mol

24. Identify the FINAL step to follow when preparing a solution.
    a) Mass out the solute and add it to the flask.
    b) Add more solvent until you reach the required amount.
    c) Mass out the solvent and add it to the flask.
    d) Add about half the required volume of solvent to the flask.

25. You start with a solution that is 0.800 mol/L and exactly 0.0700 L. You need to prepare a 0.300 mol/L solution. What is the final volume of the solution?
    a) 3.43 L
    b) 0.026 L
    c) 0.580 L
    d) 0.187 L

26. Which method of water treatment is useful for controlling disease-causing organisms such as viruses, bacteria, and parasites?
    a) Water softening
    b) Filtration
    c) Chlorination
    d) Distillation

Organic Chemistry (20 marks)

27. Which of these statements does NOT accurately describe tar sands?
    a) Tar sands provide a synthetically produced source of hydrocarbons.
    b) Tar sands are a combination of clay, sand, water, and bitumen.
    c) Tar sands can be mined and processed to extract the oil-rich bitumen.
    d) Bitumen requires no further refining and can be pumped from the ground in its natural state.

28. Which type of bond will carbon commonly form?
   a) Covalent
   b) Ionic
   c) Metallic
   d) None of these

29. All of these hydrocarbons are *unsaturated* except for
   a) Benzene
   b) Alkenes
   c) Alkanes
   d) Alkynes

30. Which of the following is the correct condensed structural formula for *butane*?
   a) $CH_3(CH_2)_3CH_3$
   b) $CH_3(CH_2)_2CH_3$
   c) $(CH_3)_3CH_3$
   d) C_4H_{10}

31. Name the following alkane: C_7H_{16}
   a) Heptane
   b) Hexane
   c) Decane
   d) Octane

32. The correct structural formula for *2,2–dimethylhexane* is:

a)
$$\begin{array}{rcl} & CH_2CH_3 & \\ & | & \\ CH_3- & C & -CH_2-CH_2-CH_2-CH_3 \\ & | & \\ & CH_3 & \end{array}$$

b)
$$\begin{array}{rcl} & CH_3 & \\ & | & \\ CH_3- & CH & -CH_2-CH_2-CH_2-CH_2-CH_3 \\ & | & \\ & CH_3 & \end{array}$$

c)
$$\begin{array}{cl} CH_3 & \\ | & \\ CH & -CH_2-CH_2-CH_2-CH_2-CH_3 \\ | & \\ CH_3 & \end{array}$$

d)
$$\begin{array}{rcl} & CH_3 & \\ & | & \\ CH_3- & C & -CH_2-CH_2-CH_2-CH_3 \\ & | & \\ & CH_3 & \end{array}$$

33. Molecules that have the same molecular formula but different structural formulas are called

a) Allotropes

b) Sterioisomers

c) Structural isomers

d) Isotopes

34. Which straight-chain alkane is a structural isomer of *3-propylheptane*?

a) 10 carbon atoms = decane

b) 6 carbon atoms = hexane

c) 9 carbon atoms = nonane

d) 5 carbon atoms = pentane

35. The correct name for the alkene $CH_3CH{=}CH_2$ is

a) propene
b) prop–2–ene
c) 2–propene
d) prop–3–ene

36. The correct structural formula for *2–methylbut–1–ene* is

a)
$$\begin{array}{ccccccc} & & & & CH_3 & & \\ & & & & | & & \\ CH_3 & - & CH & - & CH & = & CH_3 \end{array}$$

b)
$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_2 & = & C & - & CH_2 & - & CH_3 \end{array}$$

c)
$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_3 & - & C & = & CH & - & CH_3 \end{array}$$

d)
$$\begin{array}{ccccccc} & & & & & & CH_3 \\ & & & & & & | \\ CH_3 & - & CH & = & CH_2 & - & CH_2 \end{array}$$

37. The correct name for the alkyne $CH{\equiv}CCH_2CH_2CH_3$ is

a) pent–1–yne
b) 1–pentyne
c) pentyne
d) pent–4–yne

38. The correct structural formula for *3–ethylpent–1–yne* is

a)
$$\begin{array}{l} \qquad\quad CH_2CH_3 \\ \qquad\quad | \\ C \equiv C - C - CH_2 - CH_3 \end{array}$$

b)
$$\begin{array}{l} \qquad\; CH_3 \\ \qquad\; | \\ CH \equiv C - CH - CH_2 - CH_3 \end{array}$$

c)
$$\begin{array}{l} CH_2CH_3 \\ | \\ CH_2 - CH_2 - C \equiv C - CH_3 \end{array}$$

d)
$$\begin{array}{l} \qquad\qquad CH_2CH_3 \\ \qquad\qquad | \\ CH \equiv C - CH - CH_2 - CH_3 \end{array}$$

39. Give the IUPAC name for the alcohol $CH_3CH(OH)CH_2CH(CH_3)CH_3$.

a) 2–methylpentan–4–ol
b) 2–methyl–4–pentanol
c) 4–methylpentan–2–ol
d) 4–methyl–2–pentanol

40. Identify which of the following alcohols is most likely used to make hand soap.

a) Isopropyl
b) Glycerol
c) Ethanol
d) Methanol

41. Name the following carboxylic acid: $CH_3CH_2CH(CH_3)CH_2COOH$

a) 3–methylpentanoic acid
b) 2–methylbutanoic acid
c) methyl–3–pentanoic acid
d) methyl–2–butanoic acid

42 Identify which of the following carboxylic acids is responsible for the sting in ant bites.

a) Benzoic acid
b) Lactic acid
c) Formic acid
d) Acetic acid

43. Name the ester $CH_3CH_2CH_2CH_2COOCH_3$.

a) methyl butanoate
b) methylpentanoic acid
c) pentyl methanoate
d) methyl pentanoate

44. A very large molecule made of many smaller repeating units is known as

a) A monomer
b) A polymer
c) An ester
d) An allotrope

45. Which of these examples of is NOT a polymer formed by an addition reaction?

a) Graphite
b) Teflon™
c) Polypropylene
d) Synthetic rubber

46. Which of these polymers is used for moulded plastics and film?

a) Polyethylene
b) Polyvinyl chloride (PVC)
c) Teflon™
d) Polypropylene

Part C: Short Answer (32 Marks)

Answer each of the questions below using the space provided. Pay attention to the number of marks that each question is worth, as this may help you decide how much information to provide for full marks. For questions that involve calculations, show your work and check your final answer for the correct number of significant figures and the appropriate unit.

Stoichiometry ***(15 marks)***

1. How many moles of $O_{2(g)}$ react with 2.4 moles of Fe in the following rusting reaction? *(2 marks)*

$$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$$

Answer:

$$2.4\ \cancel{\text{mol Fe}} \times \frac{3\ \text{mol O}_2}{4\ \cancel{\text{mol Fe}}} = 1.8\ \text{mol}$$

(1 mark for the calculation, 1 mark for the correct answer)

2. What quantity of heat is produced in the complete combustion of 60.2 g of ethane gas (C_2H_6), according to the following balanced chemical reaction? The heat of combustion of ethane is 1560 kJ/mol and its molar mass is 30.0 g/mol. *(4 marks)*

$$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(g)}$$

Answer:

$$\text{mol} = 60.2\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{30.0\ \cancel{\text{g}}}\right) = 2.01\ \text{mol}$$

$$\text{or, mol} = \left(\frac{60.2\ \cancel{\text{g}}}{30.0\ \cancel{\text{g}}/\text{mol}}\right) = 2.01\ \text{mol}$$

$$\text{energy} = 2.01\ \cancel{\text{mol}} \times \left(\frac{1560\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) = 3140\ \text{kJ}$$

(1 mark for each step of the calculation for 3 marks total, 1 mark for the correct answer)

3. How many grams of CO_2 would be produced if 45 g of $C_6H_{12}O_6$ (glucose) reacted completely with oxygen? Glucose = 180.0 g/mol; CO_2 = 44.0 g/mol. *(4 marks)*

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$$

Answer:

$$45\ \cancel{g}\ C_6H_{12}O_6 \times \frac{1\ \text{mol}}{180.0\ \cancel{g}} = 0.25\ \text{mol}\ C_6H_{12}O_6$$ *(2 marks)*

$$0.25\ \cancel{\text{mol}\ C_6H_{12}O_6} \times \frac{6\ \cancel{\text{mol}}\ CO_2}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}} \times \frac{44.0\ g}{\cancel{\text{mol}}} = 66\ g\ CO_2$$ *(2 marks)*

4. Given 5.0 moles of sulfur and 8.4 moles of oxygen gas, as well as

$$2S + 3O_2 \rightarrow 2SO_3$$

a) Identify the limiting factor and the excess reactant. *(3 marks)*

Answer:

$$\text{moles } O_2 = 5.0\ \cancel{\text{mol S}} \times \frac{3\ \text{mol}\ O_2}{2\ \cancel{\text{mol S}}} = 7.5\ \text{mol}\ O_2$$ *(2 marks)*

7.5 mol O_2 is needed to use up all of the S. You are given more than 7.5 moles of O_2, so S is the limiting factor, and O_2 is the excess reactant. *(1 mark)*

b) Calculate the moles of excess reactant that remain. *(2 marks)*

Answer:

moles O_2 remaining = initial moles – reacted moles

= 8.4 mol – 7.5 mol

= 0.9 mol O_2 excess *(2 marks)*

Solutions ***(7 marks)***

5. Describe any three properties that are true of a solution in terms of the particulate view of matter. *(3 marks)*

   Answer:

   Any three of the following properties for one mark each: Solutions are homogeneous, their particles are spread evenly throughout the solution, they have a single phase, their particles are too small to be seen, they are transparent, their particles are too small to reflect light, their components do not settle out, and their parts cannot be separated by filtration.

6. Write the equation for dissolving $Ag_2CrO_{4(s)}$ in water. *(2 marks)*

   Answer:

   $Ag_2CrO_{4(s)} \rightarrow 2Ag^+_{(aq)} + CrO_4^{2-}{}_{(aq)}$

7. What is the number of moles of solute in 0.650 L of a 0.40 mol/L solution? *(2 marks)*

Answer:

$$0.650\ \cancel{L} \times \frac{0.40\ \text{mol}}{1\ \cancel{L}} = 0.26\ \text{mol}$$

Organic Chemistry *(10 marks)*

8. Draw the structural formulas for the following hydrocarbons. *(2 marks x 3 = 6 marks)*

   a) 2,4–dimethylpentane

   Answer:

   $CH_3CH(CH_3)CH_2CH(CH_3)CH_3$

   b) hex–2–ene

   Answer:

   $CH_3CH{=}CHCH_2CH_2CH_3$

   c) 4–ethylhex–2–yne

   Answer:

   $CH_3C{\equiv}CCH(C_2H_5)CH_2CH_3$

9. Complete the following table. *(4 marks)*

	Aliphatic Hydrocarbon	**Aromatic Hydrocarbon**
Similarities	Can be saturated or unsaturated.	Unsaturated.
Differences	Does not contain a benzene ring, does not show resonance.	Contains a benzene ring, demonstrates resonance.

NOTES

Grade 11 Chemistry
Final Practice Examination

Word Bank

Use the following word bank to help you complete the "Fill-in-the-Blank" portion of your Final Examination. Note that there may be MORE terms here than you need, so read over the list carefully before choosing the terms that you want to use. You can also use certain words more than once.

alcohol(s)	dehydrogenation	hydroxyl	ratio(s)
aliphatic	depression	immiscible	R-COOH
alkyl	diamond	increase(s)	reactant
allotrope	dilution	inorganic	saturated
amorphous carbon	electronegativity	insoluble	smaller
aqueous	electrons	isomers	soluble
aromatics	electrostatic	less	solute
benzene	emulsion	limiting	solution
boiling	endothermic	miscible	solvation
branched-chain	equal	molarity	solvent
buckminsterfullerene	esterification	moles	stoichiometry
carbon	excess	more	substituent
carboxyl	exothermic	non-polar	supersaturated
carboxylic	graphite	number	suspensions
coefficients	greater	organic	theoretical
colligative	higher	percent	unsaturated
concentration	hydrogenation	phenyl	vapour
cracking	hydrated	polar	water
crude	hydrocarbon	products	
decrease(s)	hydrogen	proportion(s)	

Grade 11 Chemistry
Final Practice Examination

Bubble Sheet

Name: ______________________________ ____________ / 46

For each Multiple Choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the exam.

	A	B	C	D		A	B	C	D		A	B	C	D		A	B	C	D
1.	❍	❍	❍	❍	14.	❍	❍	❍	❍	27.	❍	❍	❍	❍	40.	❍	❍	❍	❍
2.	❍	❍	❍	❍	15.	❍	❍	❍	❍	28.	❍	❍	❍	❍	41.	❍	❍	❍	❍
3.	❍	❍	❍	❍	16.	❍	❍	❍	❍	29.	❍	❍	❍	❍	42.	❍	❍	❍	❍
4.	❍	❍	❍	❍	17.	❍	❍	❍	❍	30.	❍	❍	❍	❍	43.	❍	❍	❍	❍
5.	❍	❍	❍	❍	18.	❍	❍	❍	❍	31.	❍	❍	❍	❍	44.	❍	❍	❍	❍
6.	❍	❍	❍	❍	19.	❍	❍	❍	❍	32.	❍	❍	❍	❍	45.	❍	❍	❍	❍
7.	❍	❍	❍	❍	20.	❍	❍	❍	❍	33.	❍	❍	❍	❍	46.	❍	❍	❍	❍
8.	❍	❍	❍	❍	21.	❍	❍	❍	❍	34.	❍	❍	❍	❍					
9.	❍	❍	❍	❍	22.	❍	❍	❍	❍	35.	❍	❍	❍	❍					
10.	❍	❍	❍	❍	23.	❍	❍	❍	❍	36.	❍	❍	❍	❍					
11.	❍	❍	❍	❍	24.	❍	❍	❍	❍	37.	❍	❍	❍	❍					
12.	❍	❍	❍	❍	25.	❍	❍	❍	❍	38.	❍	❍	❍	❍					
13.	❍	❍	❍	❍	26.	❍	❍	❍	❍	39.	❍	❍	❍	❍					

Periodic Table of the Elements

Atomic Number → 19
Symbol → **K**
Name → Potassium
Relative Atomic Mass → 39.1

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Unuhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Electronegativities Table

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 **Lanthanide Series**	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 **Actinide Series**	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Alphabetical Listing of the Elements and Their Atomic Masses

Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Actinium	(227)	Gold	197.0	Praseodymium	140.9
Aluminum	27.0	Hafnium	178.5	Promethium	(145)
Americium	(243)	Hassium	(265)	Protactinum	(231)
Antimony	121.7	Helium	4.0	Radium	(226)
Argon	39.9	Holmium	164.9	Radon	(222)
Arsenic	74.9	Hydrogen	1.0	Rhenium	186.2
Astatine	(210)	Indium	114.8	Rhodium	102.9
Barium	137.3	Iodine	126.9	Rubidium	85.5
Berkelium	(247)	Iridium	192.2	Ruthenium	101.1
Beryllium	9.0	Iron	55.8	Rutherfordium	(261)
Bismuth	209.0	Krypton	83.8	Samarium	150.4
Bohrium	(264)	Lanthanum	138.9	Scandium	45.0
Boron	10.8	Lawrencium	(257)	Seaborgium	(263)
Bromine	79.9	Lead	207.2	Selenium	79.0
Cadmium	112.4	Lithium	6.9	Silicon	28.1
Calcium	40.1	Lutetium	175.0	Silver	107.9
Californium	(251)	Magnesium	24.3	Sodium	23.0
Carbon	12.0	Manganese	54.9	Strontium	87.6
Cerium	140.1	Meitnerium	(266)	Sulfur	32.1
Cesium	132.9	Mendelevium	(256)	Tantalum	180.9
Chlorine	35.5	Mercury	200.6	Technetium	(98)
Chromium	52.0	Molybdenum	95.9	Tellurium	127.6
Cobalt	58.9	Neodymium	144.2	Terbium	158.9
Copernicium	(277)	Neon	20.2	Thallium	204.4
Copper	63.5	Neptunium	(237)	Thorium	232.0
Curium	(247)	Nickel	58.7	Thulium	168.9
Dubnium	(262)	Niobium	92.9	Tin	118.7
Dysprosium	162.5	Nitrogen	14.0	Titanium	47.9
Einsteinium	(254)	Nobelium	(259)	Tungsten	183.8
Erbium	167.3	Osmium	190.2	Uranium	238.0
Europium	152.0	Oxygen	16.0	Vanadium	50.9
Fermium	(257)	Palladium	106.4	Xenon	131.3
Fluorine	19.0	Phosphorus	31.0	Ytterbium	173.0
Francium	(223)	Platinum	195.1	Yttrium	88.9
Gadolinium	157.2	Plutonium	(244)	Zinc	65.4
Gallium	69.7	Polonium	(209)	Zirconium	91.2
Germanium	72.6	Potassium	39.1		

Names, Formulas, and Charges of Common Ions

Positive Ions (Cations)

Name	Symbol	Name	Symbol
aluminum	Al^{3+}	magnesium	Mg^{2+}
ammonium	NH_4^+	**manganese(II)**	**Mn^{2+}**
barium	Ba^{2+}	manganese(IV)	Mn^{4+}
cadmium	Cd^{2+}	mercury(I)	Hg_2^{2+}
calcium	Ca^{2+}	mercury(II)	Hg^{2+}
chromium(II)	Cr^{2+}	**nickel(II)**	**Ni^{2+}**
chromium(III)	**Cr^{3+}**	nickel(III)	Ni^{3+}
copper(I)	Cu^+	potassium	K^+
copper(II)	**Cu^{2+}**	silver	Ag^+
hydrogen	H^+	sodium	Na^+
iron(II)	Fe^{2+}	strontium	Sr^{2+}
iron(III)	**Fe^{3+}**	tin(II)	Sn^{2+}
lead(II)	**Pb^{2+}**	**tin(IV)**	**Sn^{4+}**
lead(IV)	Pb^{4+}	zinc	Zn^{2+}
lithium	Li^+		

continued

Negative Ions (Anions)

Name	Symbol	Name	Symbol
acetate	$C_2H_3O_2^-(CH_3COO^-)$	nitrate	NO_3^-
azide	N_3^-	nitride	N^{3-}
bromide	Br^-	nitrite	NO_2^-
bromate	BrO_3^-	oxalate	$C_2O_4^{2-}$
carbonate	CO_3^{2-}	hydrogen oxalate	$HC_2O_4^-$
hydride	H^-	oxide	O^{2-}
hydrogen carbonate or bicarbonate	HCO_3^-	perchlorate	ClO_4^-
chlorate	ClO_3^-	permanganate	MnO_4^-
chloride	Cl^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}
chromate	CrO_4^{2-}	dihydrogen phosphate	$H_2PO_4^-$
citrate	$C_6H_5O_7^{3-}$	silicate	SiO_3^{2-}
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfate	HSO_4^-
fluoride	F^-	sulfide	S^{2-}
hydroxide	OH^-	hydrogen sulfide	HS^-
hypochlorite	ClO^-	sulfite	SO_3^{2-}
iodide	I^-	hydrogen sulfite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4^+ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

continued

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^-	Bromate	OH^-	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^-	Bromide	IO_3^-	Iodate	O^{2-}	Oxide
ClO_3^-	Chlorate	I^-	Iodide	O_2^{2-}	Peroxide
Cl^-	Chloride	NO_3^-	Nitrate	SO_4^{2-}	Sulfate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulfide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulfite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocyanate	**3^- charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite