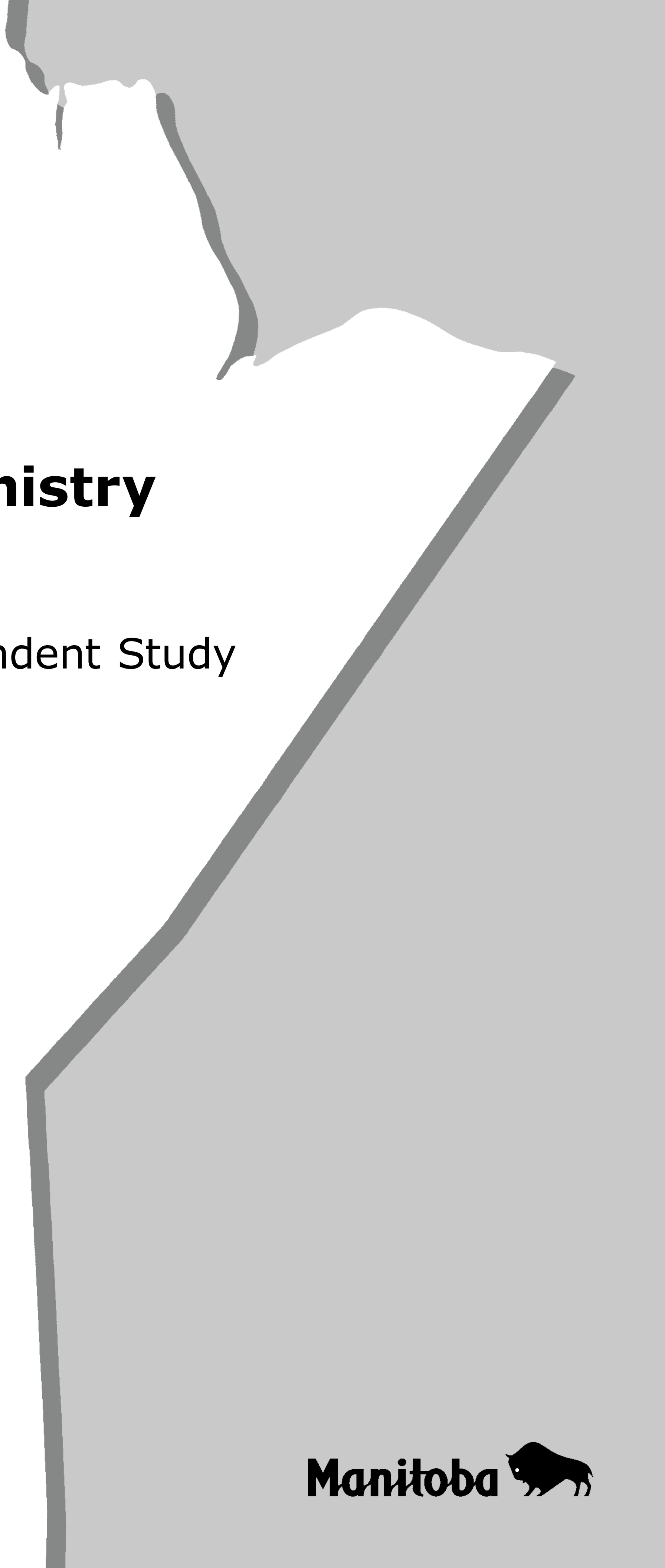

Grade 12 Chemistry (40S)

A Course for Independent Study

Field Validation Version

Manitoba

GRADE 12 CHEMISTRY (40S)

A Course for Independent Study

Field Validation Version

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This resource was published in 2013 and updated in 2019.

Available in alternate formats upon request.

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Module 5: Acids and Bases

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INTRODUCTION

Overview

Welcome to Grade 12 Chemistry!

Chemistry is the study of the interactions between matter and energy. This is most likely your second full course in chemistry. Many of the topics that are explored in Grade 11 Chemistry are built upon in this course.

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/process work 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 understand the world around you and how things work. Everything is made up of chemicals and chemistry helps to explain fireworks, osteoporosis, tooth decay, turbo chargers in engines, batteries, and many other things. This course will enable you to better your understanding and knowledge of important topics related to chemistry and daily life, and make better choices in your everyday life.

Everyone uses chemistry. Firefighters, artists, nurses, truck drivers, doctors, dentists, plumbers, pharmacists, environmentalists, physiotherapists, hairdressers, chefs, and veterinarians learn about chemistry and use that knowledge in their jobs. In fact, you will find that Grades 11 and 12 Chemistry are a requirement for admittance to many university and college programs.

How is This Course Organized?

This course consists of the following six modules, plus appendices:

- Module 1: Aqueous Reactions
- Module 2: Atomic Structure
- Module 3: Kinetics
- Module 4: Equilibrium
- Module 5: Acids and Bases
- Module 6: Electrochemistry
- 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.

> ***A note about SLO numbers***
>
> In each lesson introduction, you will find SLO numbers (e.g., SLO 1.1.1, SLO 1.1.2). These numbers are for teacher use only. If you are a student, you can disregard these numbers.

What Resources Will You Need for This Course?

Required Resources

You do not need a textbook for this course. All the content is provided directly within the course. You will, however, need access to a variety of resources.

- **Graphing or scientific calculator:** Calculator software/applications can be downloaded onto computers or electronic devices instead of using the recommended calculator.
- **A computer with Internet access:** You will require Internet access to view the Grade 12 Chemistry Module 2 and Module 5 coloured graphics available in the learning management system (LMS). If a hardcopy is required, please contact the ISO office and request the Grade 12 Chemistry Module 2 and Module 5 CD. If you do not have access to the Internet to view the coloured graphics, please contact the Distance Learning Unit at 1-800-465-9915.
- **Internet access to view animations and videos:** In some lessons, URLs are listed that link to animations and videos on the Internet. Please note that URLs often change, and although they were active at the time the course was written, it is possible that they will not be active when you try to access them. In such cases, use a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/ to find similar sites.

Some of the URLs are from the T.I.G.E.R. (Teachers' International Graphics Educational Resource) site from the North Carolina School of Science and Mathematics (NCSSM) and start with www.dlt.ncssm.edu.

The videos are also available on youtube.com. To access them, you can do one of the following:

1. Go to the YouTube NCSSM Distance Education website at www.youtube.com/user/NCSSMDistanceEd, and then type in the name of the video on the Search Channel box at the top right-hand side.

   or

2. Go to www.youtube.com and search for the video by name.

If you have difficulty finding a video mentioned in the course, you can use a search engine such as www.google.ca and search for it by name. If that doesn't work, contact your tutor/marker.

If you have access to a computer but not the Internet, you can still view the NCSSM animations and videos by requesting the CD from the Distance Learning Unit at 1-800-465-9915. The CD is titled *Chemistry Files: Animations & Videos, North Carolina School of Science and Mathematics.* Please note that the CD contains only the animations from NCSSM; it does not contain the animations from other sources.

- **Computer that can run FLV and SWF files:** Whether you view the animations and videos through the Internet or on CD, your computer must be able to run FLV and SWF files.

  There are a couple of free downloads available to allow you to view either a FLV or SWF file:

  - **Adobe Flash player:** You can download a free player at http://get.adobe.com/flashplayer/?promoid=DINNK.
  - **FLV player:** You can download a free FLV player at http://download.cnet.com/FLV-Player/3000-13632_4-10467081.html.

Terms of Use of NCSSM Animations and Videos

In order to view the animations and videos produced by the North Carolina School of Science and Mathematics (NCSSM), you must abide by the following terms:

- "You will acknowledge that these animations and videos are the property of NCSSM and will give credit to NCSSM when using them."
- "You will not profit from these animations and videos."
- "You will use these animations and videos for educational purposes only."
- "You will refrain from modifying these animations and videos."

- "You will adhere to the following statement of non-discrimination in the use of these videos and animations: In administering its affairs, the North Carolina School of Science and Mathematics is committed to equality of opportunity. It is the policy of the School to be fair and impartial in all its relations with its students, employees, and applicants for employment and to not discriminate against any person on the basis of race, colour, creed, national origin, sex, sexual orientation, religion, disability, age, or honourable service in the armed forces of the United States."

Internet Safety

If you choose to use the Internet to do research, be safe. The Internet is a valuable source of information and should be used responsibly. Talk to your parents/guardians about Internet safety, and use the following guidelines when going online:

- Choose a user name that does not tell your name, gender, age, or other personal details.
- Never give anyone private information.
- Do not answer emails from strangers.
- If someone asks you to keep your relationship with them a secret, stop talking to the person and immediately tell your parent/guardian.
- Do not email or post pictures or files.

The above is **not** a complete list because no list can possibly cover all dangerous situations. Use your common sense and be careful.

Optional Resources

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

- **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.
- **Graph paper:** Several activities will require a graph to properly analyze data. The graph paper can be purchased at any dollar store.
- **Effervescent tablets (Module 3):** Another one of the learning activities requires you to collect data using effervescent tablets (for example, Alka-Seltzer tablets). A stopwatch would also be useful for this activity.

- **Red cabbage (Module 5):** One of the learning activities will require you to boil red cabbage juice to test the pH range (acidity) of different household substances. This can be purchased at any grocery store.
- **Galvanized nails (Module 6):** One of the learning activities will require you to use a galvanized nail, a copper penny, an LED light, and a conducting wire and a lemon to construct a battery. A multimeter can also be used.

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 your course: your tutor/marker and your learning partner.

Your Tutor/Marker

Tutor/markers are experienced educators who tutor independent students and mark assignments and examinations. When you are having difficulty with something in this course, be sure to 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. Your tutor/marker information is also available 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 course work, 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 examinations with your learning partner. If you and your learning partner are taking the same course, however, your assignment work should not be identical.

Plagiarism

Plagiarism IS a big deal with serious consequences, so it’s important that you understand what it is and how to avoid it.

What is plagiarism?

In brief, plagiarism is taking someone’s ideas or words and presenting them as if they are your own.

How can you avoid plagiarism?

- Begin early. Research takes time. Allow enough time to search for, evaluate, and read sources, and to get help if you need it. Always document your sources immediately.
- Present your research by quoting and paraphrasing.
  - When you use a quote, you use the exact same words with quotation marks, and you indicate exactly where it came from.
  - When you paraphrase, you rewrite an author’s idea using your own words and you do not use quotation marks (but you also make sure to state clearly whose idea it is).
- Learn how to use different citation styles.
- Give credit where credit is due. Never pretend someone else’s idea is your own.

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 examination(s) successfully. Many of the questions on the examination(s) will be similar to the questions in the learning activities. Remember that you **will not submit learning activities to the Distance Learning Unit**.

Assignments

Each module in this course contains assignments, which you will complete and submit to the Distance Learning Unit for assessment. The assignments are worth a total of 50 percent of your final course mark.

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

The 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 mark in this course. You will write the midterm examination when you have completed Module 3.
- The **final examination** is based on Modules 4 to 6 and is worth 30 percent of your final mark in this course. 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.

To do well on each examination, you should review all the work you have completed from the modules, including all learning activities and assignments.

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 Examination(s)

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 Independent Study Option (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
  P.O. 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 this course, but the pace at which you proceed is up to you. Read the following charts for suggestions on how to pace yourself.

Chart A: Semester 1

If you want to start the 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 compete 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 compete 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.

Note

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 six times. The following chart shows you exactly what assignments you will be submitting at the end of each module.

Submission of Assignments	
Submission	Assignments You Will Submit
1	**Module 1: Reactions in Aqueous Solutions** Module 1 Cover Sheet Assignment 1.1: Solubility and Precipitation Assignment 1.2: Predicting Precipitate Formation Assignment 1.3: Writing Balanced Neutralization Reactions Assignment 1.4: Stoichiometry of Neutralization Assignment 1.5: Neutralization Calculations Assignment 1.6: Oxidation Numbers Assignment 1.7: Interpreting Redox Reactions Assignment 1.8: Balancing Redox Reactions Assignment 1.9: Redox Applications
2	**Module 2: Atomic Structure** Module 2 Cover Sheet Assignment 2.1: Applications of Line Spectra Assignment 2.2: Electron Configuration Assignment 2.3: Valence Electrons Assignment 2.4: Periodic Trends
3	**Module 3: Chemical Kinetics** Module 3 Cover Sheet Assignment 3.1: Reaction Rate Assignment 3.2: Collision Theory and Reaction Rate Assignment 3.3: Potential Energy Diagrams Assignment 3.4: Reaction Mechanisms Assignment 3.5: Rate and Order
4	**Module 4: Chemical Equilibrium** Module 4 Cover Sheet Assignment 4.1: Dynamic Equilibrium Assignment 4.2: Writing Equilibrium Law Assignment 4.3: Solving K_{eq} Problems Assignment 4.4: Le Châtelier's Principle Assignment 4.5: Interpreting a Concentration versus Time Graph Assignment 4.6: The Haber Process Assignment 4.7: Working with Solubility Product Constant Assignment 4.8: Kidney Stones

continued

Submission of Assignments (continued)	
Submission	Assignments You Will Submit
5	**Module 5: Acids and Bases** Module 5 Cover Sheet Assignment 5.1: Acid-Base Theories Assignment 5.2: Identifying Conjugate Pairs Assignment 5.3: Solving K_W Problems Assignment 5.4: Solving pH Problems Assignment 5.5: Acid and Base Equilibrium Assignment 5.6: pH and Percent Dissociation Assignment 5.7: Titration Curves Assignment 5.8: Acid-Base Predictions
6	**Module 6: Electrochemistry** Module 6 Cover Sheet Assignment 6.1: Predicting Spontaneity Assignment 6.2: History of the Voltaic Cell Assignment 6.3: The Functioning Voltaic Cell Assignment 6.4: Calculating Standard Reduction Potentials Assignment 6.5: Electroplating Assignment 6.6: Using Faraday's Law

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 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
P.O. 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?

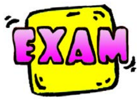
Guide graphics are used throughout this course to identify and guide you in specific tasks. Each graphic has a specific purpose, as described below.

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 the Learning Activity Answer Keys found at the end of the applicable module.

Assignment: Complete an assignment. You will submit your completed assignments to the Distance Learning Unit for assessment. You will be submitting your assignments at the end of every module.

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

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



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

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

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

Laboratory Activity: Complete an experiment at this time.

Phone or Email: Telephone or email your tutor/marker.

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!

GRADE 12 CHEMISTRY (40S)

Module 1 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 1 Assignments	**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: Solubility and Precipitation	______ /8	______ /8
☐ Assignment 1.2: Predicting Precipitate Formation	______ /15	______ /15
☐ Assignment 1.3: Writing Balanced Neutralization Reactions	______ /10	______ /10
☐ Assignment 1.4: Stoichiometry of Neutralization	______ /10	______ /10
☐ Assignment 1.5: Neutralization Calculations	______ /20	______ /20
☐ Assignment 1.6: Oxidation Numbers	______ /26	______ /26

continued

Grade 12 Chemistry (40S)

Module 1 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 1 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 1.7: Interpreting Redox Reactions	_______ /10	_______ /10
☐ Assignment 1.8: Balancing Redox Reactions	_______ /30	_______ /30
	Total: ___ **/129**	**Total:** ___ **/129**

For Tutor/Marker Use
Remarks:

GRADE 12 CHEMISTRY (40S)

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: Applications of Line Spectra	______ /10	______ /10
☐ Assignment 2.2: Electron Configuration	______ /12	______ /12
☐ Assignment 2.3: Valence Electrons	______ /10	______ /10
☐ Assignment 2.4: Periodic Trends	______ /14	______ /14
	Total: ____ **/46**	**Total:** ____ **/46**

For Tutor/Marker Use
Remarks:

GRADE 12 CHEMISTRY (40S)

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	**Mailing Address**
Distance Learning Unit 555 Main Street Winkler MB R6W 1C4	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: Reaction Rate	______ /10	______ /10
☐ Assignment 3.2: Collision Theory and Reaction Rate	______ /10	______ /10
☐ Assignment 3.3: Potential Energy Diagrams	______ /10	______ /10
☐ Assignment 3.4: Reaction Mechanisms	______ /17	______ /17
☐ Assignment 3.5: Rate and Order	______ /15	______ /15
	Total: ___ **/62**	**Total:** ___ **/62**

For Tutor/Marker Use
Remarks:

GRADE 12 CHEMISTRY (40S)

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:

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 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: Dynamic Equilibrium	_______ /5	_______ /5
☐ Assignment 4.2: Writing Equilibrium Law	_______ /10	_______ /10
☐ Assignment 4.3: Solving K_{eq} Problems	_______ /15	_______ /15
☐ Assignment 4.4: Le Châtelier’s Principle	_______ /14	_______ /14
☐ Assignment 4.5: Interpreting a Concentration versus Time Graph	_______ /10	_______ /10
☐ Assignment 4.6: The Haber Process	_______ /16	_______ /16

continued

GRADE 12 CHEMISTRY (40S)

Module 4 Cover Sheet *(continued)*

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Winkler MB R6W 4B8

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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 *(continued)*	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	____________ Date Received	____________ Date Received
☐ Assignment 4.7: Working with Solubility Product Constant	_______ /20	_______ /20
☐ Assignment 4.8: Kidney Stones	________ /5	________ /5
	Total: ___ **/95**	**Total:** ___ **/95**

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Remarks:

Grade 12 Chemistry (40S)

Module 5 Cover Sheet

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City/Town: ______________________________ Postal Code: ____________

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School Name: ____

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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: Acid-Base Theories	______ /5	______ /5
☐ Assignment 5.2: Identifying Conjugate Pairs	______ /14	______ /14
☐ Assignment 5.3: Solving K_W Problems	______ /10	______ /10
☐ Assignment 5.4: Solving pH Problems	______ /23	______ /23
☐ Assignment 5.5: Acid and Base Equilibrium	______ /10	______ /10
☐ Assignment 5.6: pH and Percent Dissociation	______ /15	______ /15

continued

Grade 12 Chemistry (40S)

Module 5 Cover Sheet *(continued)*

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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 5.7: Titration Curves	______ /16	______ /16
☐ Assignment 5.8: Acid-Base Predictions	______ /5	______ /5
	Total: ___ **/98**	**Total:** ___ **/98**
For Tutor/Marker Use		
Remarks:		

GRADE 12 CHEMISTRY (40S)

Module 6 Cover Sheet

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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	**Attempt 1**	**Attempt 2**
Which of the following are completed and enclosed? Please check (✓) all applicable boxes below.	______________ Date Received	______________ Date Received
☐ Assignment 6.1: Predicting Spontaneity	________ /13	________ /13
☐ Assignment 6.2: History of the Voltaic Cell	________ /12	________ /12
☐ Assignment 6.3: The Functioning Voltaic Cell	________ /11	________ /11
☐ Assignment 6.4: Calculating Standard Reduction Potentials	________ /15	________ /15
☐ Assignment 6.5: Electroplating	________ /14	________ /14
☐ Assignment 6.6: Using Faraday’s Law	________ /14	________ /14

continued

GRADE 12 CHEMISTRY (40S)

Module 6 Cover Sheet *(continued)*

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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 1.9: Researching Redox Reactions	________ /24	________ /24
	Total: ____ **/103**	**Total:** ____ **/103**

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Remarks:

Released 2019

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GRADE 12 CHEMISTRY (40S)

Module 1
Reactions in Aqueous Solutions

MODULE 1:
REACTIONS IN AQUEOUS SOLUTIONS

Introduction to Module 1

Water is essential to life on Earth. More than two-thirds of Earth's surface is water. In terms of weight, the cells in our bodies contain between 65% and 90% water. Many of life's reactions occur in water: they are aqueous. In this module, you will study three types of aqueous reactions: precipitation reactions, neutralization reactions, and oxidation-reduction reactions.

General Notes

Throughout this Grade 12 Chemistry course, you may notice the following:

- Assignments and learning activities are numbered sequentially. This means that their numbers will not always be the same as the lesson number.
- In some lessons, there are multiple learning activities.
- Not every lesson has both an assignment and a learning activity.
- The time indicated on the first page of each lesson is only a **guideline**. In other words, the lesson may take more or less time.

Note

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.

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	Solubility and Precipitation
2	Assignment 1.2	Predicting Precipitate Formation
3	Assignment 1.3	Writing Balanced Neutralization Reactions
4	Assignment 1.4	Stoichiometry of Neutralization
5	Assignment 1.5	Neutralization Calculations
6	Assignment 1.6	Oxidation Numbers
7	Assignment 1.7	Interpreting Redox Reactions
8	Assignment 1.8	Balancing Redox Reactions

LESSON 1: WHAT IS SOLUBILITY? (2 HOURS)

Lesson Focus

SLO C12-1-01: Explain examples of solubility and precipitation at the particulate and symbolic levels.

Lesson Introduction

There are two key terms you must know in order to understand this lesson and the ones that follow. Those terms are **solubility** and **precipitation**. In order to fully understand these concepts, you must know what occurs at the particulate level when two substances mix together. In other words, if you could see the particles, what would they be doing when forming a solution?

Solutions Terminology

Let's assume that everyone needs a brief review of Grade 11 Chemistry before moving forward. In Grade 11, you studied the dissolving process. The following paragraph summarizes important terms from Grade 11 Chemistry that will come up again in Grade 12 Chemistry.

A **solution** is a homogenous mixture composed of solvent and solute particles. A **solvent** is the dissolving medium and, in many cases, is water. The solvent is usually the substance present in the greater amount in a solution. A **solute** is the dissolved particle and is usually the substance present in a lesser quantity in a solution. In the picture on this page, the effervescent tablet is the solute, while the water is the solvent.

Solvation, also called **dissolution**, occurs when a solute dissolves in a solvent. It is the process by which solvent particles surround and dissolve solute particles. The nature of the particles determines whether solvation occurs and, ultimately, whether a solution will form. A substance is considered **soluble** if it will dissolve in a specific solvent. A substance is considered to be **insoluble** if it dissolves only very little. (Later in the course, we will quantify what this statement means.) However, note that all solutes are soluble to some degree. **Solubility** describes the amount of solute that will dissolve in a given quantity of solvent at a given temperature.

Properties of Solutions

In Grade 11 Chemistry, you learned about the properties of solutions. Assuming some review would be helpful, the following is an overview of those properties:

- Solutions are homogenous. (They look the same throughout and the particles of solute and solvent are evenly mixed.)
- Solutions are stable. (The solute will not settle out over time.)
- Both the solute and solvent pass through a filter when filtered. (In other words, you cannot easily separate the solute and solvent.)
- Solvents and solutes may be gases, liquids, or solids, as well as combinations of these three.
- Solutions are considered to be a single phase even though the components may have been in different phases before the solution was formed.

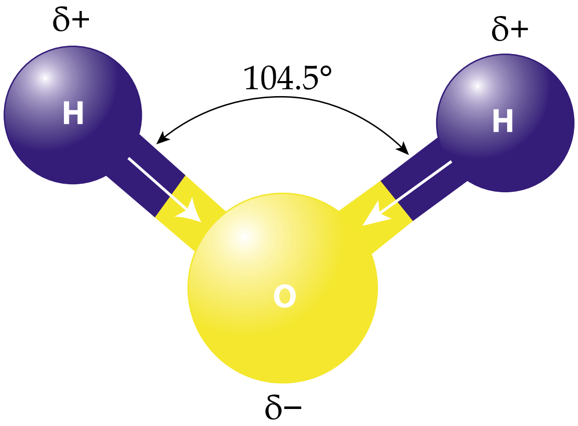
Solubility

In Grade 11 Chemistry, you learned that particles are in constant random motion. As these reactant particles collide with one another, they may form products. There are certain conditions which can favour these collisions, such as an increase in temperature, often creating more product.

Factors that affect solubility include:

1. **Nature of the solute and solvent:** You may remember that "like dissolves like." This means that, in general,
   a. polar and ionic solutes tend to be more soluble in polar solvents
   b. non-polar solutes tend to be more soluble in non-polar solvents

   Remember that water is considered to be a polar molecule due to its characteristic bent shape (shown below), which makes one side of a water molecule more positive and the other side more negative.

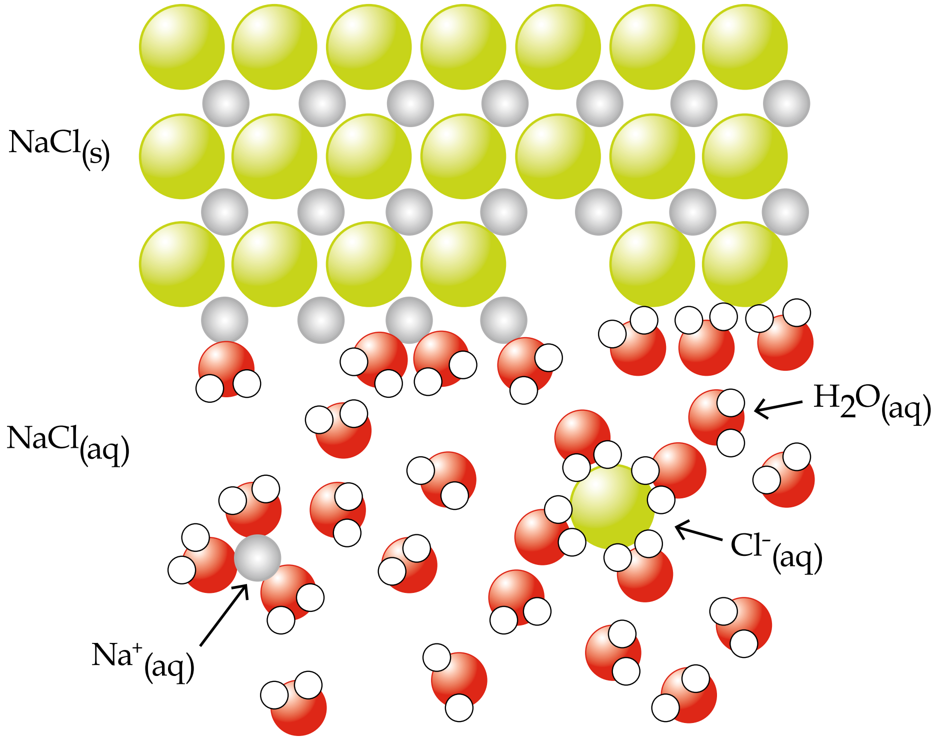


2. **Temperature:** As temperature increases, the solubility of most solids in a liquid increases. This is due to the increased kinetic energy of particles. For most gases, a decrease in temperature results in an increased solubility in liquids.
3. **Pressure:** While changes in pressure do not have a big effect on the solubility of solids and liquids, the same is not true of gases. In Grade 11 Chemistry, you learned that the solubility of a gas is directly proportional to the pressure of that gas above a liquid. For example, opening a can of carbonated soft drink reduces the gas pressure in the can, and dissolved carbon dioxide gas bubbles out of a solution.

Solution Process of Ionic Compounds

Ionic compounds consist of charged ions held together by electrostatic forces of attraction. Common examples of an ionic substance are sodium chloride (NaCl) and silver chloride (AgCl). Ionic compounds are generally solid, crystalline structures at room temperature. That being said, their solubility in water varies greatly. While NaCl (table salt) is quite soluble in water, AgCl is only slightly soluble. What accounts for this difference in solubility?

Let's examine the following *molecular level representation* of sodium chloride being dissolved by water. NaCl is the solute and water (H_2O) is the solvent.



In order for a solution to form, the solvent (water) must overcome the force of attraction that exists between the ions and break the crystal structure. The process of surrounding the solute ions with solvent molecules is called **solvation** or **dissolution**. To accomplish this, the polar water molecules orientate (arrange) themselves around each exposed ion in the solid in such a way that the following occurs:

- The **positive** end of the water molecule is attracted to the **negatively** charged chloride ion. The force of attraction pulls the chloride ion away from the crystal surface where water molecules surround it—the chloride ion is now dissolved.
- The **negative** end of the water molecule is attracted to the **positively** charged sodium ion. The force of attraction pulls the sodium ion away from the crystal surface where water molecules surround it—the sodium ion is now dissolved.

If the solute is **soluble**, the attraction between the solvent and solute particles is greater than the attraction within the solute crystal. Next, the solute ions are pulled into the solvent and become completely surrounded by the solvent molecules. As a result, salt is broken down into *individual* cations (Na^+) and anions (Cl^-), shown is the diagram above.

This separation of ions is called **dissociation**. When ionic compounds dissolve in water, the ions separate (or dissociate) from each other and mix with the water molecules. NaCl will readily dissolve in water to form a solution of sodium ions (Na^+) and chloride ions (Cl^-), mixed with water molecules. This dissociation can be expressed as follows:

$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

You may wonder why water is not written in the chemical equation. In all **aqueous** reactions, water is the solvent; therefore, it is understood to be present.

If you have access to the Internet, you can watch an animation that shows NaCl being dissolved by water online at www.mhhe.com/physsci/chemistry/essentialchemistry/flash/molvie1.swf.

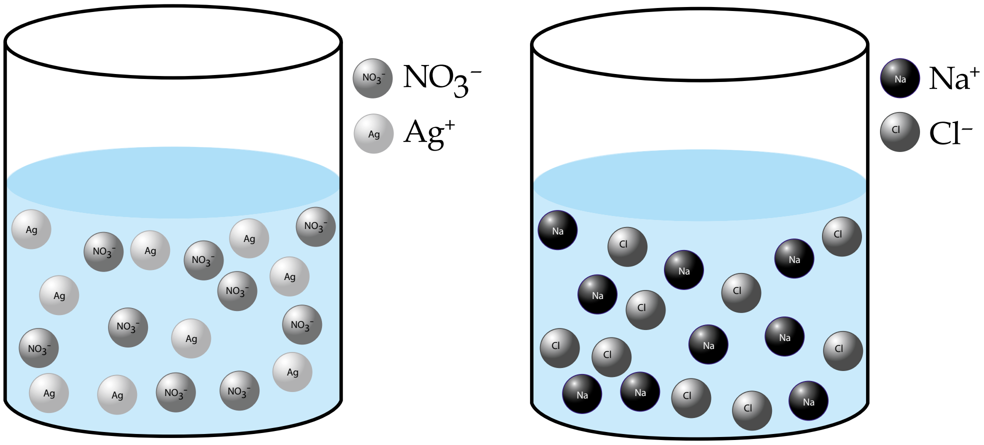
If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Precipitation Reactions

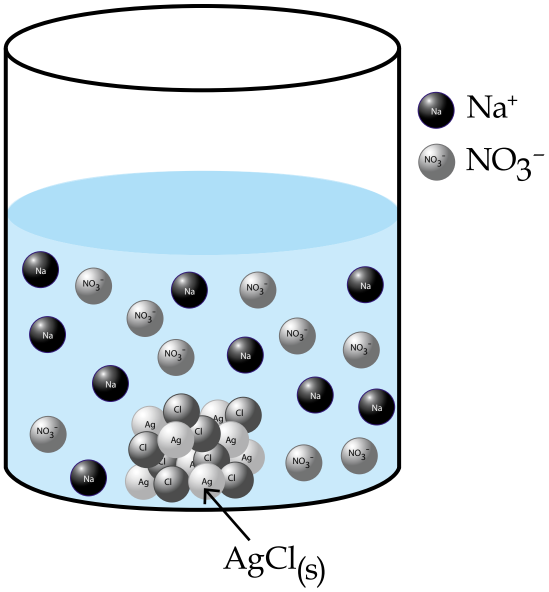
When two solutions are mixed during an aqueous reaction, a substance with low solubility sometimes forms. This substance forms a solid that separates from the solution and settles to the bottom. This is because the attraction between the ions is greater than the attraction between water and the individual ions. Consequently, the ions will not stay in a solution. These are called **precipitation reactions**. Note that not all mixtures of ions create a precipitation reaction.

Example 1

When a solution of aqueous silver nitrate ($AgNO_3$) and aqueous sodium chloride (NaCl) are mixed, the two clear, colourless solutions form a cloudy mixture. The cloudy mixture clears as a white solid settles to the bottom of the container. Before mixing these two solutions, they would appear as homogeneous mixtures, as shown below.



When the two solutions are mixed together, the force of attraction between the silver (Ag^+) and chloride (Cl^-) ions is stronger than the force keeping them in a solution. When the two ions come into contact, they form the compound silver chloride (AgCl). Eventually, the number of ions joining together becomes so large that the water cannot suspend them and they settle to the bottom as a solid. The process in which ions leave a solution and form an ionic solid is called **precipitation**.



Notice from the diagram above that the sodium and nitrate ions remain in a solution. Since the ions separate in a solution, they are allowed to react separately. The sodium ions and nitrate ions do not participate in the reaction; they just "stand around and watch." Ions that do not participate in a precipitation reaction and still remain moving freely in a solution are called **spectator ions**. In this example, the sodium and nitrate ions are the spectator ions.

Example 2

When solutions of lead (II) nitrate and potassium iodide are prepared, they will dissociate as shown below.

$$Pb(NO_3)_{2(aq)} \rightarrow Pb^{2+}{}_{(aq)} + 2NO_3{}^{-}{}_{(aq)}$$

$$KI_{(aq)} \rightarrow K^{+}{}_{(aq)} + I^{-}{}_{(aq)}$$

Therefore, when the two solutions are mixed, the following ions will be present:

Pb^{2+}, $NO_3{}^{-}$, K^{+}, I^{-}

Two new compounds will form: lead iodide (PbI_2) and potassium nitrate (KNO_3). PbI_2 happens to be only slightly soluble in water; therefore, it will precipitate. The balanced chemical equation below shows that the lead iodide (PbI_2) formed is a solid, which indicates that it will precipitate.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

If you have access to the Internet, you can watch the animation that shows the reaction between these two soluble ionic compounds, forming a slightly soluble precipitate at www.crescent.edu.sg/crezlab/webpages/PptReaction3.htm.

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

To help you practice what you have just learned about solubility and precipitation, you will now complete Learning Activity 1.1. This learning activity, and all others in the course, will help you prepare for your assignments and examinations. You will not submit learning activities to the Distance Learning Unit for assessment. You will, however, submit all assignments to the Distance Learning Unit.

Learning Activity 1.1: Introduction to Solutions

Answer each of the questions on solubility and precipitation in the space provided.

1. List two substances that are purchased as solids but are then made into solutions before use.
2. List some examples of solutions that you use in everyday life.
3. Explain why you think clothes might be easier to clean in hot water than in cold water.
4. Complete the dissociation equations for each of these ionic compounds.
    a. $NaBr_{(s)} \rightarrow$
    b. $HCl_{(aq)} \rightarrow$
    c. $Fe(OH)_{3(s)} \rightarrow$
    d. $Na_2(SO_4)_{(aq)} \rightarrow$

Check the answer key.

After reading the lesson summary below, complete Assignment 1.1. It will help you review the key concepts about solubility and precipitation. This assignment (along with all other assignments) is **worth marks**. **All assignments in Module 1 will be submitted together at the end of the module,** except for Assignment 1.9, which will be submitted at the end of Module 6.

Lesson Summary

In this lesson, you learned that a **precipitate** is a solid that separates from a solution. A **precipitation reaction** is one in which the aqueous (dissolved) ions combine to form an insoluble solid. In the next lesson, you will learn to predict when precipitates will form.

Assignment 1.1: Solubility and Precipitation (8 marks)

Answer each of the questions on solubility and precipitation in the space provided.

1. Complete the dissociation equations for each of these ionic compounds. *(4 marks)*
   a. $NaF_{(s)} \rightarrow$
   b. $Na_2S_{(s)} \rightarrow$
   c. $HI_{(s)} \rightarrow$
   d. $MgCl_{2(s)} \rightarrow$

2. Describe the main events that occur when an ionic compound (such as NaCl) dissolves in water. *(4 marks)*

Notes

LESSON 2: PREDICTING PRECIPITATE FORMATION (2 HOURS)

Lesson Focus

SLO C12-1-03: Use a table of solubility rules to predict the formation of a precipitate.

Lesson Introduction

So far in this course you have expanded on your knowledge of solubility and precipitation. You know that not every solid will dissolve in liquid. You also learned that when two dissolved solids are mixed together in an aqueous medium, sometimes a precipitate will form. How can you predict which solids are soluble and which ones are not? How do you know when to expect precipitate formation in a chemical reaction? In this lesson, you will use a table of solubility rules to predict the formation of a precipitate.

Solubility Rules

In Lesson 1, you learned that ionic compounds, such as salts, differ in their solubility in water. Most **insoluble** salts, such as AgBr, will dissolve to some extent in water, if only to a negligible degree. For this reason, we will more accurately use the term **low solubility** for such compounds.

Several ionic substances have a low solubility in water because the force of attraction between the ions is greater than the force water molecules can exert to dissociate the ions. These compounds do not dissolve in water and, therefore, will settle to the bottom. There are some generally accepted rules that can be used to determine if a substance is soluble in water.

Following is a table that summarizes the solubility of ionic compounds in water. Note that a full-size version of this table will also appear in the appendices section at the end of this course. (See Appendix E: Solubility Chart.)

Table of Solubility Rules

Ion	Soluble with...	Low Solubility with...
alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	essentially all	
fluoride ion (F^-)	all others	Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Li^+, Pb^+
hydrogen ion (H^+)	essentially all	
ammonium ion (NH_4^+)	essentially all	
nitrate (NO_3^-)	essentially all	
acetate (CH_3COO^-) ($C_2H_3O_2^-$)	essentially all	Ag^+
chloride (Cl^-) bromide (Br^-) iodide (I^-)	all others	Ag^+, Pb^{2+}, Hg^{2+}, Hg^+, Cu^+, Tl^+
sulphate (SO_4^{2-})	all others	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}, Ag^+
sulphide (S^{2-})	all alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	all others
hydroxide (OH^-)	all alkali ions, H^+, NH_4^+, Sr^{2+}, Ba^{2+}	all others
phosphate (PO_4^{3-}) carbonate (CO_3^{2-}) sulphite (SO_3^{2-})	all alkali ions, H^+, NH_4^+	all others
chromate (CrO_4^{2-})	all others	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+

These rules are a set of generalizations about the solubility of ionic compounds in water at 25°C in concentrations of 0.1 mol/L. As such, these rules may vary if these conditions change. It is important to note that the rules apply to **compounds**. Individual **ions** cannot be soluble or insoluble, and positive ions must always be associated with negative ions in a compound.

Predicting Solubility

In the previous section, several rules were outlined that you can use to predict if a compound is soluble or not. Note that the term "low solubility" is used rather than "insoluble" since a small amount (even if it is *very* small) does actually dissolve. Remember that these are all aqueous solutions, meaning that the solvent is always water. For this reason, water is not included in the chemical reaction.

When using the chart (Appendix E) to predict solubility, first find the negative ion (left column), then find its partner positive ion (central column) and, finally, find the solubility (right column). Read the table carefully and be especially aware that species such as Cu^+ and Cu^{2+} are different, as are NO_3^- and NO_2^-. Also, note that on some lines, the term "all others" means that the ion is **soluble**, while on other lines it means the ion has **low solubility**.

Example 1

Use the solubility rules to predict if sodium sulphate (Na_2SO_4) is water-soluble.

Solution: Find the sulphate ion in the chart. You will see that only the specified ions have a low solubility. Sodium is not listed, therefore, it falls under "all others" which means that sodium sulphate is **soluble**.

Ion	Soluble with…	Low Solubility with…
sulfate (SO_4^{2-})	all others	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}, Ag^+

Example 2

Use the solubility rules to predict if silver nitrate ($AgNO_3$) is water-soluble.

Solution: Find the nitrate ion in the chart. You will see that nitrate ions are soluble with almost all positive ions; therefore, silver nitrate is **soluble**.

Ion	Soluble with…	Low Solubility with…
nitrate (NO_3^-)	essentially all	

Learning Activity 1.2: Using Solubility Rules

Use the Table of Solubility Rules to predict which of the following substances have low solubility and which are soluble.

1. KBr ___________________________________
2. $NaNO_3$ ___________________________________
3. $BaSO_4$ ___________________________________
4. $FeCl_2$ ___________________________________
5. tin (II) sulphate ___________________________________
6. lithium carbonate ___________________________________
7. silver iodide ___________________________________
8. ammonium bromide ___________________________________
9. copper (II) chloride ___________________________________
10. calcium sulphide ___________________________________

Check the answer key.

Precipitation Reactions

A **precipitation reaction** is one in which aqueous (dissolved) ions from two compounds undergo a double replacement reaction and, if any of the new products are insoluble, a solid drops out of solution. The solid formed when two soluble compounds produce an insoluble product is referred to as a **precipitate**. In other words, if a negative ion (anion) and a positive ion (cation) combine to form an ionic compound with low solubility, precipitation will occur. Precipitation reactions can be predicted by consulting the solubility table for each of the possible products.

Using your solubility chart, you can watch the following animation online at www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s2_3.swf and see if you can predict if precipitation will occur between three different compounds.

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Use the following steps to predict the formation of a precipitate:

1. Write out the **molecular equation**.

This will include the compounds in the original solutions as reactants and the products of a double replacement.

2. Consult the solubility table to determine if the products are soluble or have low solubility.
3. Write the **total ionic equation**.

This is similar to the molecular equation except that the reactants and any soluble products are written as dissolved ions and the insoluble products are written in molecular form. It helps if you write out the dissociation reactions for each of your reactants and soluble products first.

4. Write out the **net ionic equation**.

Cancel out all ions that appear on both sides of the total ionic equation. These are called spectator ions because they are not involved in a reaction.

Example 1

Dilute solutions of the salt $AgNO_3$ (solution 1) and NaCl (solution 2) are mixed. Determine if there are any precipitates and write the net ionic reaction.

1. Start with a balanced equation, with unknown states for the products:

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(?)} + NaNO_{3(?)}$$

2. Determine the solubility of the products:

Checking the solubility table, we find that AgCl has low solubility; $NaNO_3$ is soluble. This means that AgCl will precipitate from the solution. We can now write in the state symbols for the products.

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{\mathbf{(s)}} + NaNO_{\mathbf{3(aq)}}$$

[balanced chemical equation]

3. Total ionic equation:

First, write out the dissociation reactions for your reactants and **soluble** products.

$$AgNO_{3(S)} \rightarrow Ag^{+}_{(aq)} + NO_3{}^{-}_{(aq)}$$

$$NaCl_{(S)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$NaNO_{3(s)} \rightarrow Na^{+}_{(aq)} + NO_3{}^{-}_{(aq)}$$

Now, rewrite the balanced chemical equation substituting with these ions and the precipitate:

$$\mathrm{Ag^+_{(aq)} + NO_3{}^-_{(aq)} + Na^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)} + Na^+_{(aq)} + NO_3{}^-_{(aq)}}$$

[total ionic equation]

4. Net ionic equation:

Cancel out spectator ions that appear on both sides of the equation and rewrite.

$$\mathrm{Ag^+_{(aq)} + \cancel{NO_3{}^-_{(aq)}} + \cancel{Na^+_{(aq)}} + Cl^-_{(aq)} \rightarrow AgCl_{(s)} + \cancel{Na^+_{(aq)}} + \cancel{NO_3{}^-_{(aq)}}}$$

$\mathrm{Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}}$ **[net ionic equation]**

Example 2

Determine if a precipitate will form in the reaction between aqueous lead (II) nitrate and aqueous potassium iodide. Write out the balanced chemical equation, total ionic, and net ionic equations.

1. Start with a balanced chemical equation, with unknown states for the products:

   $$\mathrm{Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(?)} + 2KNO_{3(?)}}$$

2. Determine the solubility of the products using the solubility rules table. In this case, lead(II) ions have a low solubility with iodide ions. Therefore, lead(II) iodide will form a precipitate. Potassium nitrate is soluble according to the solubility rules table. Complete the balanced chemical equation by adding in the state symbols for the products:

   $$\mathrm{Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}}$$

   [balanced chemical equation]

3. Write out the total ionic equation.

Start by writing out the dissociation reactions for all soluble (aqueous) compounds in the reaction.

$$\mathrm{Pb(NO_3)_{2(aq)} \rightarrow Pb^{2+}_{(aq)} + 2NO_3{}^-_{(aq)}}$$

$$\mathrm{2KI_{(aq)} \rightarrow 2K^+_{(aq)} + 2I^-_{(aq)}}$$

$$\mathrm{2KNO_{3(aq)} \rightarrow 2K^+_{(aq)} + 2NO_3{}^-_{(aq)}}$$

Substitute the ions of these soluble compounds into the balanced chemical equation.

$$\mathrm{Pb^{2+}_{(aq)} + 2NO_3{}^-_{(aq)} + 2K^+_{(aq)} + 2I^-_{(aq)} \rightarrow PbI_{2(s)} + 2K^+_{(aq)} + 2NO_3{}^-_{(aq)}}$$

[total ionic equation]

4. Cancel out spectator ions and write out the net ionic equation.

$$Pb^{2+}_{(aq)} + \cancel{2NO_3^-{}_{(aq)}} + \cancel{2K^+_{(aq)}} + 2I^-_{(aq)} \rightarrow PbI_{2(s)} + \cancel{2K^+_{(aq)}} + \cancel{2NO_3^-{}_{(aq)}}$$

$Pb^{2+}_{(aq)} + 2I^-_{(aq)} \rightarrow PbI_{2(s)}$ **[net ionic equation]**

Learning Activity 1.3: Predicting Precipitation

Write the complete set of equations (balanced chemical, total ionic, and net ionic) for each of the following reactions. Then, determine whether or not a precipitate will form in each case.

1. $MgCl_{2(aq)} + Na_3PO_{4(aq)}$
2. $Al_2(SO_4)_{3(aq)}$ and $Ba(NO_3)_{2(aq)}$
3. $(NH_4)_2S_{(aq)}$ and $Co(NO_3)_{2(aq)}$

Check the answer key.

Lesson Summary

In this lesson, you learned how to predict when a compound would precipitate. In other words, you learned when its ions would not stay in a solution. You used the Table of Solubility Rules to help you make these predictions. In the next lesson, you will learn about another type of aqueous reaction, called **neutralization**.

Notes

Assignment 1.2: Predicting Precipitate Formation (15 marks)

1. Use the Table of Solubility Rules to predict whether the following compounds are soluble or whether they have low solubility. *(5 marks)*
   a. aluminum hydroxide ($Al(OH)_3$) ________________________________
   b. lead (II) chloride ($PbCl_2$) ______________________________________
   c. nickel (II) iodide (NiI_2) _______________________________________
   d. barium phosphate ($Ba_3(PO_4)_2$) _________________________________
   e. copper (I) bromide (CuBr) _______________________________________

2. Use the solubility rules to predict whether a precipitate will form when the following pairs of aqueous solutions are mixed. For those cases in which a precipitate forms, write the net ionic equation for the reaction.
   (3 x 2 marks = 6 marks)
   a. $H_2SO_{4(aq)}$ and $Pb(NO_3)_{2(aq)}$

   b. $Na_3PO_{4(aq)}$ and $FeCl_{3(aq)}$

   c. $Pb(C_2H_3O_2)_{2(aq)}$ and $LiCl_{(aq)}$

continued

Assignment 1.2: Predicting Precipitate Formation (continued)

3. What are spectator ions? *(1 mark)* Identify the spectator ions in each of the reactions in Question 2. *(3 marks)*

LESSON 3: NEUTRALIZATION REACTIONS

Lesson Focus

SLO C12-1-04: Write balanced neutralization reactions involving strong acids and bases.

Lesson Introduction

In the previous lesson, you learned that not all compounds stay in solution. If a compound dissociates and its ions do not stay in solution, a solid will form and settle out upon standing. In this lesson, you will continue your study of aqueous reactions while focusing on neutralization reactions. What is a strong acid? What is a strong base? What occurs when the two are mixed? You will discover the answers to these and other questions in this lesson.

How Are Acids Named?

In this as well as some of the following lessons, the names of acids will be used. Generally, acids are formed when non-metal ions combine with hydrogen. Understanding how acids are named will help you identify them in a reaction as well as write their names from a formula. There are two types of acids that you will encounter in this module:

- simple (also called **binary**)
- polyatomic

Naming Acids with Simple Anions (Binary Acids)

To name a binary acid, follow these three steps:

Step 1: Name the ionic compound as you normally would.

Step 2: Remove the ending (-ite, -ide, or -ate) and add the suffix “ic” to the root word.

Step 3: Add the prefix “hydro” and the word acid.

Example

Name: HCl

Step 1: hydrogen chloride

Step 2: hydrogen chloric

Step 3: hydrochloric acid

Naming Acids with Polyatomic Anions

The process for naming polyatomic acids follows a different set of rules. In many cases, the steps used for the binary acids will work with the following exceptions:

- The suffix "ic" is added to the root word if the anion name ends in "ate."
- The suffix "ous" is added to the root word if the anion name ends in "ite."

Examples

HNO_3: There is a H+ ion and a nitrate ion. The acid formed is named **nitric acid**.

H_2SO_4: There are two H+ ions and a sulphate ion. The acid formed is named **sulphuric acid**.

Neutralization Reactions

A neutralization reaction is a type of double replacement reaction between a strong acid and a strong base, where a salt and water are formed as products. The word "neutralization" is appropriate because the acid and base properties of H^+ and OH^- are neutralized. The resulting solution, having no excess hydrogen or hydroxide ions, is neither an acid nor a base—it is neutral.

As you just read, the products of a neutralization reaction are a salt and water. The H^+ ions (from the acid) and OH^- ions (from the base) combine to form water molecules. The salt produced in an acid-base reaction is not strictly limited to common table salt (although that is an example of a salt). A **salt** is any ionic, crystalline compound which can be produced from the neutralization of a strong acid and a strong base.

General Equation

$$\underset{\text{acid}}{HX_{(aq)}} + \underset{\text{base}}{YOH_{(aq)}} \rightarrow \underset{\text{water}}{H_2O_{(l)}} + \underset{\text{salt}}{YX_{(aq\ or\ s)}}$$

The water is formed from the H^+ ion of the acid (H*X*) and the OH^- ion of the base (*YOH*).

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

The negative ions of the acid will unite with the positive ions of the base to form a **salt**.

$$Y^+_{(aq)} + X^-_{(aq)} \rightarrow YX_{(aq\ or\ s)}$$

Example: $H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow Na_2SO_{4\ (aq)} + 2\ H_2O_{(l)}$

Using the Table of Solubility Rules you used in the last lesson, you can confirm whether each product will be aqueous or solid. Make sure the equation is balanced!

If you have difficulty predicting the products of a neutralization reaction, you can write the **total ionic equation**, showing all ions that are in a solution. You can then cancel the **spectator ions** and complete the **net ionic equation**, like you did in Lesson 2. For this example, it would be as follows:

Total ionic equation:
$2\ H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\ OH^-_{(aq)} + 2\ Na^+_{(aq)} \rightarrow 2\ Na^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\ H_2O_{(l)}$

Net ionic equation: $2\ H^+_{(aq)} + 2\ OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}$

Reduce coefficients as needed: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Note

For future neutralization reactions, use only the balanced chemical equations.

Strong Acids and Bases

In this module, only strong acids and strong bases will be referenced. What is a strong acid? What is a strong base? A **strong acid** will dissociate completely (100%) into ions. For example, if there are 100 molecules of HCl dissolved in water, 100 ions of H^+ and 100 ions of Cl^- are produced. For an animation of how a strong acid ionizes, go to the following link and click on the HCl tab.

www.dlt.ncssm.edu/tiger/Flash/electrochem/StrongAcidIonization.html

You can also find this animation at www.dlt.ncssm.edu/tiger/chem6.htm. Look for the file titled "StrongAcidIonization.html" or "StrongAcidIonization.exe."

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

If you do not have access to the Internet, but have access to a computer, contact the Independent Study Option office and request the CD titled *Chemistry Files: Animations & Videos, North Carolina School of Science and Mathematics*. Please note that the CD contains only the animations from NCSSM; it does not contain the animations from sources other than NCSSM.

For more information on the URLs, including those from NCSSM, please refer to the introduction of this course.

The following are six strong acids with which you should be familiar:

- hydrochloric acid (HCl)
- hydrobromic acid (HBr)
- hydroiodic acid (HI)
- sulfuric acid (H_2SO_4)
- nitric acid (HNO_3)
- perchloric acid ($HClO_4$)

A **strong base** also completely dissociates into ions. For example, if there were 100 molecules of NaOH dissolved into water, 100 Na^+ ions and 100 OH^- ions would be produced. For a demonstration of how a strong base ionizes, go to the following link and click on the NaOH tab.

www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/acm2s2_1.swf

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Here are some examples of strong bases:

- sodium hydroxide (NaOH)
- potassium hydroxide (KOH)
- rubidium hydroxide (RbOH)
- calcium hydroxide ($Ca(OH)_2$)*
- barium hydroxide ($Ba(OH)_2$)*

*These bases are completely soluble in solutions of 0.01 mol/L or less.

Learning Activity 1.4: Neutralization Reactions

1. Complete and balance the following neutralization reactions. Be sure to include state symbols.
    a. $HBr_{(aq)} + NaOH_{(aq)} \rightarrow$
    b. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow$
    c. $H_2SO_{4(aq)} + KOH_{(aq)} \rightarrow$
    d. $NaOH_{(aq)} + HI_{(aq)} \rightarrow$
    e. $Ca(OH)_{2(aq)} + HCl_{(aq)} \rightarrow$
    f. $HNO_{3(aq)} + KOH_{(aq)} \rightarrow$
2. What is a **strong base**? Give an example.
3. What is a **salt**? Give an example.

Check the answer key.

Lesson Summary

In this lesson, you learned that a neutralization reaction is a double replacement reaction between a strong acid and a strong base. Neutralization is another type of aqueous reaction, one that produces a salt and water as the end result. At the end of this lesson, you practiced how to write balanced neutralization reactions. In the next lesson, you will perform a laboratory activity to demonstrate the stoichiometry of a neutralization reaction between a strong base and a strong acid.

Notes

Assignment 1.3: Writing Balanced Neutralization Reactions (10 marks)

1. Write balanced neutralization reactions given the following reactants. Be sure to include state symbols. *(4 marks)*

   a. $HCl_{(aq)} + Mg(OH)_{2(aq)} \rightarrow$

   b. $H_2SO_{4(aq)} + Ca(OH)_{2(aq)} \rightarrow$

   c. $HClO_{4(aq)} + KOH_{(aq)} \rightarrow$

   d. $HNO_{3(aq)} + NaOH_{(aq)} \rightarrow$

2. What happens in a neutralization reaction? Give an example to support your explanation. *(4 marks)*

continued

Assignment 1.3: Writing Balanced Neutralization Reactions (continued)

3. What is a strong acid? Give an example. *(2 marks)*

LESSON 4: STOICHIOMETRY OF NEUTRALIZATION (2 HOURS)

Lesson Focus

SLO C12-1-05: Perform a laboratory activity to demonstrate the stoichiometry of a neutralization reaction between a strong base and a strong acid.

Lesson Introduction

In the previous lesson, you learned about neutralization, another type of aqueous reaction. In this lesson, you will use that knowledge and perform a laboratory activity to demonstrate the stoichiometry of a neutralization reaction between a strong base and a strong acid.

Concentrated and Dilute versus Weak and Strong

Does concentrated mean the same thing as strong? Does dilute mean the same thing as weak? The answer to these questions, as you will soon see, is that they do not mean the exact same thing. Generally speaking, the terms **concentrated** and **dilute** refer to the number of **moles of hydroxide** (OH^-) or **hydrogen ions** (H^+) dissolved in a given volume of solution. For the purposes of this course, the solvent will always be water.

Example: 4 mol/L is more concentrated than 0.04 mol/L, as there are more moles in the same volume of solution.

The terms **strong** and **weak** refer to the extent of ionization of an acid or base.

In other words, these terms relate to how easily **molecules of acids or bases dissociate into ions**.

Recall that a **strong acid**, such as HCl, dissociates 100% into ions.

And so, a concentrated solution of HCl has a relatively high number of moles dissolved in a given volume of water. Note that regardless of volume, whether you have 25 mL or 5 litres, it is still a concentrated acid. If the same sample of HCl is added to a large volume of water, the solution becomes **dilute**.

Likewise, a strong base dissociates 100% into ions. Consequently, a concentrated solution of NaOH has a relatively high number of moles dissolved in a given volume of water. Again, regardless of volume, it is still a concentrated base whether you have a small or large volume of it. If the same sample of NaOH is added to a large volume of water, the solution becomes **dilute**. This is not the same thing as a **weak base** (or acid). Ammonia is a weak base because a small percent of the molecules dissociate into ions. Solutions of ammonia can be either concentrated or dilute depending on the amount of ammonia dissolved in a given volume of water.

Interestingly, you can have a *weak* acid or base that is *concentrated*. In this scenario, the acid or base in question does not dissociate 100% into ions, even though there may be a high concentration of it in solution. An example of this would be 6.0 mol/L NH_3 (6 moles of a weak base dissolved in 1 litre of water). Likewise, a *strong* acid or base can be *dilute*. This would mean that there are few moles of the given acid or base dissolved in the volume of water, but what there is dissociates 100% into ions. An example of this would be 0.1 mol/L HCl (0.1 moles of a strong acid dissolved in 1 litre of water).

pH

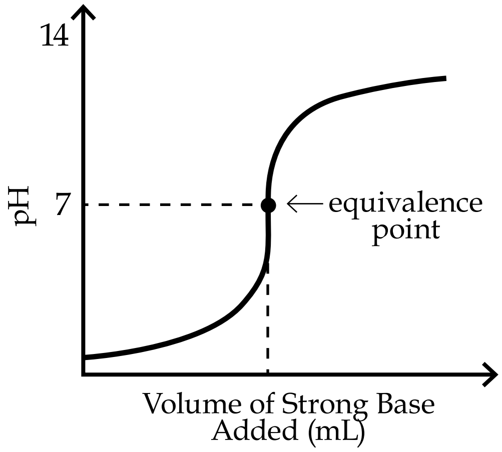
In previous science and chemistry courses, you were probably introduced to the pH scale. In Module 5 (Acids and Bases) of this course, you will be looking at this concept in depth and making pH calculations. For this module, only an understanding of what the scale represents is needed.

The pH scale is a logarithmic scale that measures the hydrogen ion concentration in acidic and basic solutions. Historically, acids were defined as substances that increase the hydrogen concentration in a solution. Pure water is considered to be neither acidic nor basic and has a pH of 7. Any solution with a pH of 7 is considered **neutral**: it is neither acidic nor basic. **Acidic solutions will have pH values below 7; basic solutions will have pH values above 7**. The pH scale extends, for practical reasons, from 0 through 14, but values outside this range are possible.

Titration

Some compounds, like sodium hydroxide, are very **hygroscopic**. In other words, they absorb significant amounts of moisture when in solid form. When mixing solutions with hygroscopic compounds, this added moisture means that mass measurements are inaccurate and the exact concentration of a solution will not be known. To determine the exact concentration of such a solution, called **standardization**, a **titration** is performed. Titration is a laboratory technique in which a carefully measured volume of an aqueous reactant of unknown concentration is slowly added to a known volume of another reactant solution with a known concentration. The known solution is often referred to as the **titrant, stock reagent** or **stock solution**. A specialized measuring tube with a valve on the bottom, called a **burette and stopcock**, is usually used for titrations because it allows for the careful, controlled addition of a reagent. In the analysis of acids and bases, a titration is conducted until the unknown solution is completely neutralized. This is called the **equivalence point**. To assist in determining when the equivalence point is reached, an indicator (a chemical that changes colour with pH) is often added. To see how a titration is set up and how it is performed, watch the video at www.youtube.com/watch?v=i7jnSaf1Muc. (If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.) There are many indicators that can be used; each with a particular pH range in which it changes colour. When the pH has reached a point where the indicator changes colour, it is called the **end point**. The indicator selected for a titration should have an end point that matches the equivalence point.

A pH meter can also be used to collect data and then to create a graph of pH versus the volume of titrant. This graph can then be used to identify the equivalence point. A strong acid when titrated with a strong base will produce a pH of 7 at the equivalence point. The same is true if a strong base is titrated with a strong acid. A sudden change in pH will occur when the equivalence point is reached on the graph.



Titration Calculations

The concentration of an unknown solution can be determined with the following data: balanced neutralization equation, volume of unknown solution used, volume of stock solution used to reach equivalence point, and concentration of the stock solution.

Example

A student wants to standardize a solution of sodium hydroxide she just mixed. She places 25.0 mL of the NaOH solution in a flask and adds a few drops of bromothymol blue indicator (which changes from blue in basic solution to yellow at pH 7). The student ends up adding 32.72 mL of a 0.129 mol/L stock solution of hydrochloric acid before the solution changes colour.

What was the molarity (concentration) of the sodium hydroxide?

- Balanced neutralization reaction:

  $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

  There is a 1:1 mole ratio between the acid and base for neutralization to occur.

- Using our known stock concentration and volume added, find the number of moles of HCl used in the neutralization.

  $\text{Concentration}_{(HCl)} = \text{Moles}_{(HCl)} / \text{Volume}_{(HCl)}$

Therefore,

$$\begin{aligned}\text{Moles}_{(HCl)} &= \text{Concentration}_{(HCl)} \times \text{Volume}_{(HCl)}\\ &= 0.129 \text{ mol/L} \times 0.03272 \text{ L}\\ &= 4.22 \times 10^{-3} \text{ mol HCl}\end{aligned}$$

- Using the mole ratio from the balanced equation, determine the moles of NaOH and calculate the concentration.

  At neutralization: moles NaOH = moles HCl = 4.22×10^{-3} mol

$$\begin{aligned}\text{Concentration}_{(NaOH)} &= \text{Moles}_{(NaOH)} / \text{Volume}_{(NaOH)}\\ &= 4.22 \times 10^{-3} \text{ mol NaOH} / 0.0250 \text{ L}\\ &= 0.169 \text{ mol/L}\end{aligned}$$

The concentration of the NaOH solution has been standardized at 0.169 mol/L.

Learning Activity 1.5: Titration

The video for this learning activity can be found at www.dlt.ncssm.edu/core/Chapter16-Acid-Base_Equilibria/Chapter16-Labs/st_acid_st_base_titration.htm.

You can also access it by

1. going to the NCSSM T.I.GE.R. site at www.dlt.ncssm.edu/tiger/chemvid8.htm
2. scrolling down the page until you find the video titled **Strong Acid Strong Base Titration—Acid Base Titrations Lab Part 1**

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

If you do not have access to the Internet, but have access to a computer, contact the Independent Study Option office and request the CD that contains all of the animations from www.dlt.ncssm.edu. Please note that the CD contains only the animations from www.dlt.ncssm.edu; it does not contain the animations from any other sources.

You will observe the following:

- 50.0 mL of **0.1068 mol/L HCl** (acid) is added to a flask.
- An indicator (**bromothymol blue**) that changes different colours in acids and bases is added to the flask.
- The **base** (0.200 mol/L NaOH) of known concentration (**standard solution** or **titrant**) is slowly added to the flask of acid.
- The base is added until the indicator changes the solution from yellow to blue, indicating that neutralization has occurred. (If the unknown solution is a base, the standard solution is an acid and vice versa.)
- Neutralization occurs when the indicator changes colour (yellow to blue). This is called the **end point** of titration.
- At this point, the solution has a pH of 7.
- Experimentally, equivalence was reached at 26.50 mL NaOH.

continued

Learning Activity 1.5: Titration (continued)

1. As you watch the video, pause at 4:00 and 4:18 minutes for the data given at the end of the titration and record the pH on the table below.

   By titration: Equivalence: 26.50 mL

Volume of NaOH (mL)	pH
0.00	
5.00	
10.00	
15.00	
20.00	
21.00	
22.00	
23.00	
24.00	
25.00	
26.00	
27.00	
28.00	
29.00	
30.00	
35.00	
40.00	
45.00	
50.00	

2. Write a balanced chemical equation for the reaction between HCl and NaOH.
3. Write the net ionic equation for the reaction between HCl and NaOH.

continued

Learning Activity 1.5: Titration (continued)

4. Graph the data. The pH will be on the vertical (y) axis and the volume of NaOH will be on the horizontal (x) axis. Be sure to label your axes and provide an appropriate title for your graph. Label the equivalence point on the graph.
5. What volume of base was required for neutralization to occur? Explain how you know this.
6. Calculate the number of moles of NaOH required to neutralize the sample of HCl.
7. Calculate the number of moles present in 50.0 mL of HCl in the laboratory.
8. Using the balanced molecular equation, determine the ratio of moles between the sodium hydroxide and the hydrochloric acid.
9. Using the number of moles obtained in Questions 5 and 6, determine the experimental ratio of moles between the NaOH and the HCl.

Check the answer key.

Lesson Summary

In this lesson, you learned that the process called **titration** is the addition of a known amount of a known solution to determine the concentration of another solution. This is a carefully controlled neutralization involving a strong acid and a strong base.

Notes

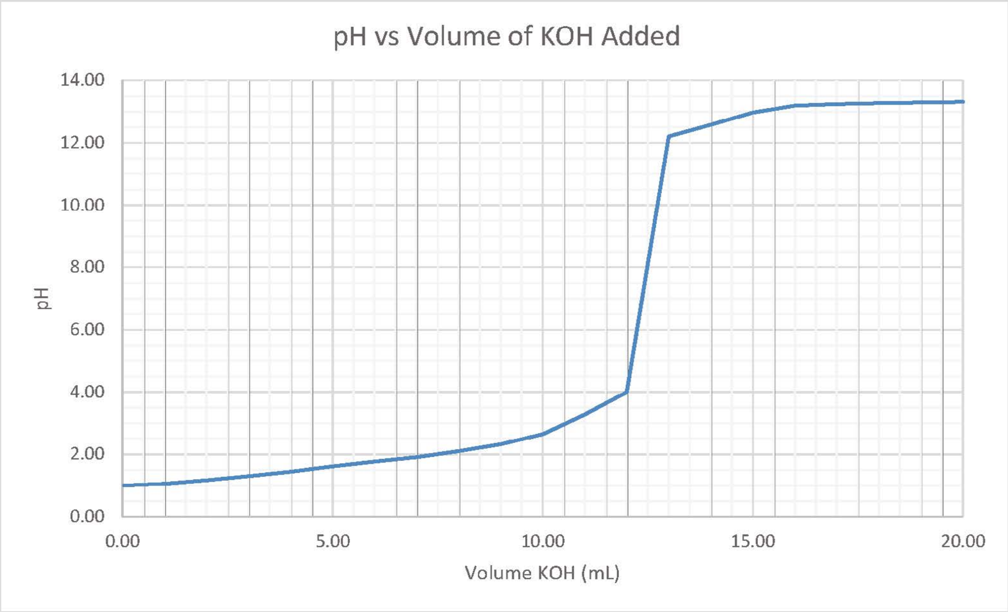
Assignment 1.4: Stoichiometry of Neutralization (10 marks)

Muhammed performed a titration between 0.150 mol/L potassium hydroxide (KOH) and 0.075 mol/L nitric acid (HNO_3). Muhammed added 25.00 ml of nitric acid (HNO_3) to an Erlenmeyer flask and then titrated the acid with the potassium hydroxide (KOH). He used a pH probe to collect the following data. Phenolphthalein indicator, which is clear and colourless in acidic solutions and pink in basic solutions, was also added to indicate a colour change when neutralization occurs.

Volume of KOH (mL)	pH
0.00	1.00
1.00	1.06
2.00	1.17
3.00	1.30
4.00	1.44
5.00	1.62
6.00	1.77
7.00	1.91
8.00	2.10
9.00	2.32
10.00	2.65
11.00	3.28
12.00	4.00
13.00	12.21
14.00	12.58
15.00	12.98
16.00	13.20
17.00	13.25
18.00	13.28
19.00	13.30
20.00	13.32

continued

Assignment 1.4: Stoichiometry of Neutralization (continued)



1. What volume of base was required for neutralization to occur? *(1 mark)*

   ____

   ____

   ____

2. Calculate the number of moles of KOH required to neutralize the sample of HNO_3. *(2 marks)*

3. Calculate the number of moles present in 25.00 mL of HNO_3 in the laboratory.
   (2 marks)

continued

Assignment 1.4: Stoichiometry of Neutralization (continued)

4. Write a balanced chemical equation for the reaction. Be sure to include state symbols. *(2 marks)*

5. What would you expect to see in the flask by the point where 13.00 mL of KOH has been added? *(1 mark)*

6. What is the ratio of H^+(aq) to OH^-(aq) at the neutralization point? *(1 mark)*

7. What has happened to the H^+ ions at the equivalence point? *(1 mark)*

Notes

LESSON 5: CALCULATING CONCENTRATION (2 HOURS)

Lesson Focus

SLO C12-1-06: Calculate the concentration or volume of an acid or a base from the concentration and volume of an acid or a base required for neutralization.

Lesson Introduction

In a previous lesson, you were introduced to terms such as *concentrated acid* and *weak base*. In this lesson, you will learn how to calculate the concentration or volume of an acid or a base from the concentration or volume of acid or base required for neutralization.

Solving Neutralization Problems

In the previous lesson, you worked with data that was produced when a certain volume of NaOH was added to neutralize a certain volume of HCl. You already knew the volume of HCl, and the concentrations of both acid and base. From the data, you were able to see the volume of NaOH required to neutralize the sample of acid and you completed stoichiometric calculations with this information.

If you know the concentration or volume of an acid or base, you can calculate the concentration or volume of acid or base required to achieve neutralization. When neutralization occurs between a base and an acid, the moles of H^+ are equivalent to the moles of OH^-:

Moles of acid/base required = Moles of acid/base used

$$\frac{\text{mol acid}}{\text{coefficient acid}} = \frac{\text{mol base}}{\text{coefficient base}}$$

Alternatively, the number of moles of base required to neutralize the acid present can be calculated by multiplying:

Moles of base (mol) = Concentration of base (mol/L) x Volume of base (in litres)

Using the number of moles for the unknown acid or base, you can then calculate its **concentration** using the following relationship:

$$C = \frac{\text{moles}}{\text{volume}}$$

Using the number of moles for the unknown acid or base, you can then calculate its **volume** using the following relationship:

$$\text{Volume} = \frac{\text{moles}}{C}$$

Recall: 1000 mL = 1.000 L

Problem-Solving Method

To gain a better understanding of the relationship between volume and concentration of acids and bases in a neutralization reaction, use the following five steps:

1. Write the **balanced chemical equation** for the neutralization reaction.
2. Under the balanced chemical equation, set up a **cvn** table to ensure that all quantities are accounted for. Fill in all the known quantities.

   c = concentration (mol/L)

   v = volume (L) (Recall that 1.000 L = 1000 mL)

   n = moles (mol)
3. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).
4. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base) using the formula mol = C x V.
5. Calculate the **required volume or concentration** of the acid or base using the formula C = mol/V.

Example 1

In the reaction of 35.0 mL of liquid drain cleaner containing NaOH, 50.08 mL of 0.409 mol/L HCl must be added to neutralize the base. What is the concentration of the base in the cleaner?

1. Write a balanced chemical equation. Include a cvn table and fill in the values given to you in the problem.

   $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

	$NaOH_{(aq)}$	$HCl_{(aq)}$
c (mol/L)	**0.586 mol/L** (from step 4)	0.409 mol/L
v (L)	0.035 L	0.05008 L
n (mol)	**0.0205 moles** (from step 2)	**0.0205 mol** (from step 3)

2. Calculate the number of moles of HCl using the formula mol = C x V

   $$\text{mol HCl} = (0.05008\ \text{L})(0.409\ \text{mol/L})$$
   $$= \mathbf{0.0205\ mol\ HCl}$$

3. Use the balanced equation for the mole ratio between HCl and NaOH, and solve for the number of moles of NaOH.

   $$\text{mol NaOH} = 0.0205\ \cancel{\text{mol HCl}} \times \frac{1\ \text{mol NaOH}}{1\ \cancel{\text{mol HCl}}}$$

   mol NaOH = **0.0205 moles NaOH**

4. Use the formula C = mol/V, to solve for the concentration of NaOH.

   $$[\text{NaOH}] = \frac{0.0205\ \text{moles}}{0.0350\ \text{L}} = \mathbf{0.586\ mol/L}$$

Once all of the cells in the cvn table are filled in, you know all the unknowns for the problem.

Example 2

Calculate the volume of 0.256 mol/L $Ba(OH)_2$ that must be added to neutralize 46.0 mL of 0.407 mol/L $HClO_4$.

1. Write a balanced equation. Be sure to include a cvn table.

   $Ba(OH)_{2(aq)} + 2HClO_{4(aq)} \rightarrow 2H_2O_{(l)} + BaCl_{2(aq)}$

	$Ba(OH)_{2(aq)}$	$2HClO_{4(aq)}$
c (mol/L)	0.256 mol/L	0.407 mol/L
v (L)	**0.0365 L** (from step 4)	0.046 L
n (mol)	**0.00935 mol** (from step 3)	**0.0187 mol** (from step 2)

2. Calculate the number of moles of $HClO_4$ using the formula mol = C x V

   $$\begin{aligned} \text{mol } HClO_4 &= (0.0460 \text{ L})(0.407 \text{ mol/L}) \\ &= \mathbf{0.0187 \text{ mol } HClO_4} \end{aligned}$$

3. Solve for the number of moles of $Ba(OH)_2$ by setting up the ratio **number of moles : coefficient** from a balanced equation.

   $$\frac{0.0187 \text{ mol } HClO_4}{2 \text{ mol } HClO_4} = \frac{\text{mol } Ba(OH)_2}{1 \text{ mol } Ba(OH)_2}$$

$$\frac{0.0187 \text{ mol } HClO_4}{2 \text{ mol } HClO_4} = \frac{\text{mol } Ba(OH)_2}{1 \text{ mol } Ba(OH)_2}$$

   0.00935 mol $Ba(OH)_2$ = mol $Ba(OH)_2$

4. Use the formula V = mol/C to solve for the volume of $Ba(OH)_2$.

   $$V = \frac{\text{mol}}{C} = \frac{0.00935 \text{ mol}}{0.256 \text{ mol/L}} = 0.0365 \text{ L}$$

   Volume of $Ba(OH)_2$ = 36.5 mL

Learning Activity 1.6: Concentration Calculations

Use the problem-solving method outlined in this lesson to perform the following calculations. Show your work and make sure your answers include the correct units and significant digits.

1. Calculate the concentration of KOH if 60 mL of this base is neutralized by 25 mL of a 0.50 mol/L solution of HCl.
   a. Write the **balanced chemical equation** for the neutralization reaction.
   b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).
   c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).
   d. Calculate the **required concentration** of the base.
2. Calculate the concentration of H_2SO_4 if 60.0 mL of this acid is neutralized by 10 mL of a 2.0 mol/L solution of NaOH.
   a. Write the **balanced chemical equation** for the neutralization reaction.
   b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).
   c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).
   d. Calculate the **required concentration** of the acid.
3. How many mL of 0.50 mol/L KOH will exactly neutralize 15.0 mL of a 2.0 mol/L solution of H_2SO_4?
   a. Write the **balanced chemical equation** for the neutralization reaction.
   b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).
   c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).
   d. Calculate the **required volume** of the base.

continued

Learning Activity 1.6: Concentration Calculations (continued)

4. How many mL of 1.0 mol/L $Ba(OH)_2$ solution must be added to 45.1 mL of a 0.20 mol/L HBr solution to make a neutral solution?
    a. Write the **balanced chemical equation** for the neutralization reaction.
    b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).
    c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).
    d. Calculate the **required volume** of base.

Check the answer key.

Lesson Summary

In this lesson, you learned a step-by-step method for solving neutralization problems. Using this method, you can calculate the concentration or volume of an unknown acid or a base from the concentration or volume of a known acid or base.

Assignment 1.5: Neutralization Calculations (20 marks)

Answer each of the following.

1. How many moles of sodium hydroxide are required to neutralize 0.20 mol of nitric acid?
    a. Write the **balanced chemical equation** for the neutralization reaction. *(2 marks)*
    b. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base). *(2 marks)*

2. How many mL of 0.45 mol/L HCl solution must be added to 25 mL of a 0.10 mol/L KOH solution to make a neutral solution?
    a. Write the **balanced chemical equation** for the neutralization reaction. *(2 marks)*

	$HCl_{(aq)}$	$KOH_{(aq)}$
c (mol/L)		
v (L)		
n (mol)		

    b. Use the concentration and volume of the known acid or base to calculate the **moles** of the **known substance** (acid or base). *(2 marks)*

continued

c. Use the coefficients from the balanced equation to determine the moles of the unknown substance (acid or base). *(2 marks)*

d. Calculate the **required volume** of the acid. *(2 marks)*

3. If 40.8 mL of 0.106 mol/L sulphuric acid neutralizes 61.8 mL of potassium hydroxide solution, what is the concentration of the base?

a. Write the **balanced chemical equation** for the neutralization reaction. *(2 marks)*

	$H_2SO_{4(aq)}$	2 $KOH_{(aq)}$
c (mol/L)		
v (L)		
n (mol)		

b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base). *(2 marks)*

c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base). *(2 marks)*

d. Calculate the **required concentration** of the base. *(2 marks)*

LESSON 6: OXIDATION AND REDUCTION (2 HOURS)

Lesson Focus

SLO C12-1-08: Outline the development of scientific understanding of oxidation and reduction reactions.
Include: gain and loss of electrons, oxidizing agent, and reducing agent

SLO C12-1-09: Determine the oxidation numbers for atoms in compounds and ions.

Lesson Introduction

What is happening to carbon in methane that is burning? Can you build a fire without oxygen? How can you explain the rusting of the bike frame that you left outside all year? In this lesson, you will be introduced to oxidation, a term that was first applied to the combining of oxygen with other elements. You will also learn about reduction, which originally meant the removal of oxygen from a compound. Finally, you will learn how to determine oxidation numbers for atoms in compounds or ions in a **redox reaction**.

Oxidation-Reduction

In the lesson introduction, the question "Can you build a fire without oxygen?" was posed. You have probably noticed that fire needs oxygen, otherwise burning will not occur. Why is this? Generally speaking, **oxidation** used to mean the combining of oxygen with other elements. Burning a substance is an example of oxidation, as oxygen from the air is added to the elements being burned.

Historically, **reduction** referred to the removal of oxygen from a compound. If reduction takes place, there will be a decrease (or reduction) in the mass of the material being reduced.

In Grade 11 Chemistry, you may have observed the burning of magnesium metal. This is an example of combustion, or oxidation, through burning:

$$2\,Mg_{(s)} + O_{2(g)} \rightarrow 2\,MgO_{(s)}$$

If written in the ionic form, it becomes:

$$2\ \mathrm{Mg}^0 + \mathrm{O}_2{}^0 \rightarrow 2\ (\mathrm{Mg}^{2+}\mathrm{O}^{2-})$$

In the above reaction, magnesium has been oxidized to MgO (magnesium oxide) by the reaction with oxygen gas. If you look at the charges, you will notice that the magnesium has gone from a 0 charge to a 2^+ charge. Meanwhile, the oxygen gas has gone from a 0 charge to a 2^- charge. Noting these changes, chemists came to redefine oxidation and reduction based on loss and gain of electrons. Therefore, an **oxidation-reduction reaction**, or **redox** (pronounced "ree-docks") **reaction**, is a reaction in which electrons are transferred (lost or gained) between reactants.

What Is an Oxidation Reaction?

Recall that oxidation reactions were first defined as the combination of oxygen with another element to produce oxides, such as the reaction forming magnesium oxide in the previous section. Redefined according to movement of electrons, **oxidation** is the process by which **electrons are removed** from an atom or ion.

If electrons are removed, then you would expect to see the charge of the atom or ion *increase* (or **become more positive**) as a result of being oxidized. Let's go back to the magnesium oxide reaction to demonstrate this concept.

$$2\ \mathrm{Mg}^0 \rightarrow \mathrm{Mg}^{2+} + 2\ e^-$$

Since magnesium went from a charge of 0 to a charge of 2^+, it tells us that two electrons were lost (or removed) in the process. Magnesium, having lost two electrons, now has two *more* protons than electrons. This charge imbalance results in the formation of a positively charged ion (a cation). Remember that elements become more positive by losing electrons, not gaining protons.

Note that, when dealing with redox reactions only, the states are unknown and, as a result, are left off.

Note

- When an atom goes from a 0 charge to a positive charge, electrons have been removed.
- When an atom or ion becomes more positively charged, oxidation has taken place.

Example 1:

Iron slowly turns to rust, according to the following reaction:

$$2\ Fe_{(s)}\ +\ 3\ O_{2(g)}\ \rightarrow\ 2\ Fe_2O_{3(s)}$$

If written in the ionic form, it becomes the following:

$$2\ Fe^0\ +\ 3\ O_2{}^0\ \rightarrow 2\ \ Fe^{3+}{}_2O^{2-}{}_3$$

In this example, iron oxidizes (gains oxygen) to form iron oxide. Since iron went from a charge of 0 to a charge of 3^+, you know that three electrons were lost (or removed). Iron, having lost three electrons, now has three *more* protons than electrons, resulting in a charge of 3^+.

Example 2:

Carbon burns in air according to the following reaction:

$$C_{(s)}\ +\ O_{2(g)}\ \rightarrow\ CO_{2(g)}$$

To help keep track of electron movement, we can write the ion charges as follows:

$$C^0\ +\ O_2{}^0\ \rightarrow\ C^{4+}O^{2-}O^{2-}$$

Carbon oxidizes (gains oxygen) to form carbon dioxide. Since carbon went from a charge of 0 to a charge of 4^+, you know that four electrons were lost (or removed). The carbon ion now has four *more* protons than electrons, resulting in a charge of 4^+.

Historically, oxidation was referred to simply as a reaction where oxygen is gained. These oxidation reactions include combustion and burning. In the above reaction, you can infer that it is an oxidation reaction by the simple fact that carbon gains oxygen in a combination (or synthesis) reaction.

Note

To help you associate oxidation with the removal of electrons, remember the acronym LEO. Each letter of the acronym represents a word that creates a helpful reminder: **L**oss of **E**lectrons is **O**xidation. Acronyms, or phrases used to remember things, are also referred to as mnemonic devices.

What Is a Reduction Reaction?

Reduction reactions were first defined as the loss of oxygen from a compound.

Redefined according to movement of electrons, **reduction** is the process by which **electrons are added** to an atom or ion. If electrons are added, then you would expect to see the charge of the atom or ion **decrease** (or **become more negative**) as a result of being oxidized. Let's refer to the magnesium oxide reaction once more.

$$2Mg^0 + O_2^0 \rightarrow 2\,(Mg^{2+}O^{2-})$$

Isolating the oxygen: $O_2^0 + 4\,e^- \rightarrow 2\,O^{2-}$

In this case, oxygen went from a charge of 0 to a charge of 2^-, indicating that two electrons were gained in the process. The oxygen ion, having gained two electrons, now has two **more** electrons than protons. This charge imbalance results in the formation of a negatively charged ion (an anion).

Note

- When an atom goes from a 0 charge to a negative charge, electrons are gained.
- When an atom or ion becomes more negatively charged, reduction has taken place (for example, a change from +3 to +2 is a gain of an electron and is, therefore, reduction).

Notice that the charge must be conserved on both sides of the equation. On the left side, there is a net charge of 4^-. On the right side there are two oxygen ions, each with a 2^- charge, for a net of 4^-. In all equations, the charges on both sides must be the same, and **the number of electrons lost must equal the number of electrons gained**.

Note

- The charge must be **conserved** on both sides of the equation.
- The number of electrons lost must equal the number of electrons gained.

Example 3:

Iron ore is reduced to its metallic iron by the removal of oxygen.

$$2\,Fe_2O_{3(s)} + 3\,C_{(s)} \rightarrow 4\,Fe_{(s)} + 3\,CO_{2(g)}$$

If written in the ionic form, it becomes

$$2\,(Fe_2^{3+}O_3^{2-}) + 3\,C^0 \rightarrow 4\,Fe^0 + 3\,(C^{4+}O^{2-}O^{2-})$$

In the reaction above, iron went from a charge of 3^+ to a charge of 0, indicating that three electrons were gained. The iron ion now has an equal number of protons and electrons, resulting in a charge of 0.

Note

To help you associate reduction with the gain of electrons, remember the acronym GER. Each letter of the acronym represents a word that creates a helpful reminder: **G**aining **E**lectrons is **R**eduction.

You might put together the two acronyms: "LEO the lion goes GER."

Learning Activity 1.7: Interpreting a Redox Reaction

1. If a neutral atom becomes positively charged, has it been oxidized or reduced? Why?
2. If an ion with a 1^- charge acquires a 2^- charge, has it been oxidized or reduced? Why?
3. Consider this equation: $Mg^0 + Co^{2+} \rightarrow Co^0 + Mg^{2+}$
   a. Which substance is oxidized?
   b. Which substance is reduced?
4. Consider this equation: $Fe^{3+} + Cu^{1+} \rightarrow Fe^{2+} + Cu^{2+}$
   a. Which substance is oxidized?
   b. Which substance is reduced?
5. Why must oxidation and reduction occur together?
6. Complete the following statements:
   In a balanced redox reaction, one substance is ________________________ when it loses electrons. In the same balanced redox reaction, the other substance is ________________________ and it therefore gains electrons.

Check the answer key.

Assigning Oxidation Numbers

Now that you can identify an oxidation and a reduction reaction, you will discover that it can sometimes be challenging to see what is being reduced and what is being oxidized. For this reason, oxidation numbers are very useful. An **oxidation number** is a positive or negative number assigned to an atom according to a set of rules. These numbers are values assigned to individual atoms within molecules or to ions, for the purpose of tracking ownership of electrons within the species. Oxidation numbers are based on the assumption that an electron belongs entirely to one atom and not to any other adjacent atom in the species.

In the previous lesson, you worked with individual charges and watched to see when the charge became more positive or more negative. You may already have noticed that the charge of an ion is written as 2^+. An oxidation number is written as +2 to keep the two notations separate.

Note

- A change in the oxidation number of an atom or ion indicates a redox reaction has occurred.
- Oxidation numbers are also used to help balance the redox reaction.

There are accepted rules for assigning oxidation numbers. The wording of these rules may vary slightly from one resource to another, but the values they outline are all the same. The following is a set of oxidation rules for you to use.

Oxidation Numbers Rules

Rule 1: The oxidation number of any elemental atom (or multiple of itself) that is not an ion is 0.

Examples: $C = 0$ $H_2 = 0$ $O_2 = 0$

Rule 2: An ion's (or complex ion's) oxidation number is its charge.

Examples: $Na^+ = +1$ $P^{3+} = +3$ $S^{2-} = -2$ $CO_3^{2-} = -2$

Rule 3: Hydrogen's oxidation number is +1, except in metal hydrides where H is the anion (e.g., CaH_2 or LiH) and it is –1.

Rule 4: The oxidation number of oxygen is –2, except in peroxides (i.e., H_2O_2, Na_2O_2) where it is –1 and in combination with fluorine (O = +2).

Rule 5: In a compound, the sum of all of the oxidation numbers of each part must equal the total charge of that compound or complex ion.

Examples:	NaCl			$CaCl_2$			SO_4^{2-}		
	Na	Cl	Overall Oxidation Number	Ca	Cl_2	Overall Oxidation Number	S	O_4	Overall Oxidation Number
	+1	-1	= 0	+2	-1 x 2 -2	= 0	+6	-2 x 4 -8	= -2

You may notice that there aren't rules for all atoms, particularly those in compounds or complex ions. This is because most atoms have more than one oxidation state. To determine the oxidation numbers for atoms in these cases, we first assign oxidation numbers for atoms for which we have rules. We then assign oxidation numbers to the unknown atoms by following rules 2 and 5.

A small table can be a useful aid for determining oxidation numbers, as you will see in the three examples that follow. In each example, you are determining the oxidation number of the element written in **bold** letters.

Example 1

$H\mathbf{N}O_3$

Rule 3 tells us that the oxidation number of H^{1+} = +1, and rule 4 tells us that O = –2. These numbers can be written above the atoms in the table as indicated.

The total oxidation number is calculated on the bottom by multiplying the oxidation number by the number of each atom present (i.e., for H, +1 x 1 atom = +1; for O, –2 x 3 atoms = –6).

+1	?	–2	The individual oxidation numbers are written on top.
H	**N**	O_3	
+1	? (+5)	–6	The total oxidation numbers are written on the bottom.

Rule 5 tells us the sum of the oxidation numbers in the bottom row must be zero because this is a neutral molecule.

Thus, the ? on the bottom line must be +5.

Since there is only one **N**, the oxidation number of **N** must be +5.

Example 2

$Na_3\mathbf{P}O_4$

Rule 2 tells us that the oxidation number of Na^{1+} = +1, and rule 4 tells us that O^{2-} = –2. These numbers can be written in the appropriate place as indicated.

The total oxidation number is calculated on the bottom (i.e., for O, –2 x 4 = –8; for Na, +1 x 3 = +3).

+1	?	–2	The individual oxidation numbers are written on top.
Na_3	**P**	O_4	
+3	? (+5)	–8	The total oxidation numbers are written on the bottom.

Rule 5 tells us the sum of the bottom oxidation numbers must be 0.

Thus, the ? on the bottom line must be +5.

Since there is only one **P** atom, the oxidation number of **P** must be +5.

Example 3

$\mathbf{Cr_2}O_7^{2-}$

This is a complex ion with an overall charge of 2–. This time, the bottom oxidation numbers must total –2. Rule 4 tells us that O = –2.

?	–2
$\mathbf{Cr_2}$	O_7
? (+12)	–14

According to rule 5, –14 + ? = –2; therefore, the ? below the chromium atom must be +12.

However, there are 2 **Cr** atoms; therefore, the oxidation number of each **Cr** must be +6.

Learning Activity 1.8: Assigning Oxidation Numbers

1. What is the oxidation number of each element in the following?
    a. Fe ______________________________
    b. Ca^{2+} ______________________________
    c. SO_2 ______________________________
    d. Na_2SO_4 ______________________________
    e. CO_3^{2-} ______________________________
    f. $MgCl_2$ ______________________________

2. In each of the following, determine the oxidation number of the element that is in **bold**.
    a. $\mathbf{Ag}$ ______________________________
    b. $\mathbf{Br}_2$ ______________________________
    c. $KN\mathbf{O}_3$ ______________________________
    d. $K\mathbf{Cl}O_3$ ______________________________
    e. $Pb\mathbf{S}O_4$ ______________________________
    f. $\mathbf{K}_3PO_4$ ______________________________
    g. $N\mathbf{O_3}^-$ ______________________________
    h. $\mathbf{N}H_4^+$ ______________________________
    i. $\mathbf{Na}_2O^2$ ______________________________
    j. $\mathbf{S}O_2$ ______________________________

Check the answer key.

Lesson Summary

In this lesson, you learned about oxidation and reduction in terms of loss and gain of electrons. To help you associate oxidation with the loss of electrons, you were introduced to the acronym LEO. To help you remember that reduction is the gain of electrons, you were introduced to the acronym GER. You also learned that an **oxidation number** is a positive or negative number assigned to an atom according to a set of rules. You used these rules to determine the oxidation number of various elements found in compounds. In the next lesson, you will learn about the oxidizing agent, reducing agent, oxidized substance, and reduced substance. Knowing these concepts will help you to identify and describe reactions as redox or non-redox.

Assignment 1.6: Oxidation Numbers (26 marks)

1. Complete the following Oxidation-Reduction Compare and Contrast Frame. *(10 marks)*

Oxidation	**Reduction**
Historical Definition *(1 mark)* ____ ____ ____ Example *(1 mark)*	Historical Definition *(1 mark)* ____ ____ ____ Example *(1 mark)*
Present Definition *(1 mark)* ____ ____ ____ Example *(1 mark)*	Present Definition *(1 mark)* ____ ____ ____ Example *(1 mark)*
Mnemonic Device *(1 mark)*	Mnemonic Device *(1 mark)*

continued

Assignment 1.6: Oxidation Numbers (continued)

2. Determine the oxidation number of each element in these formulas.

   a. P_2O_5 ____ *(2 marks)*

   b. NH_4Cl ____ *(3 marks)*

   c. $Na_2Cr_2O_7$ ____ *(3 marks)*

   d. $Ba(OH)_2$ ____ *(3 marks)*

   e. $Ca(NO_3)_2$ ____ *(3 marks)*

3. Describe how an oxidation number changes in each of these cases.

   a. An atom gains an electron. *(1 mark)*

   ____

   ____

   ____

   ____

   ____

   ____

   ____

   b. An atom loses an electron. *(1 mark)*

   ____

   ____

   ____

   ____

   ____

   ____

   ____

LESSON 7: IDENTIFYING REDOX REACTIONS (2 HOURS)

Lesson Focus

SLO C12-1-10: Identify and describe reactions as redox or non-redox.
Include: oxidizing agent, reducing agent, oxidized substance, and reduced substance

Lesson Introduction

In the previous lesson, you were introduced to oxidation numbers as useful tools to help you see where electrons are lost and gained. In this lesson, you will continue working with oxidation-reduction reactions. In addition to identifying redox reactions by electrons lost and gained, you will learn to describe the oxidizing agent, the reducing agent, the oxidized substance, and the reduced substance.

Identifying Redox Reactions

Oxidation and reduction always accompany one another to form a complete reaction. In the previous lesson, you learned that a **redox** reaction is one where

electrons gained = electrons lost

In other words, the electrons are lost and gained at the same time so that the number of electrons lost (oxidation) must equal the number gained (reduction).

Note

Oxidation = loss of electrons—**LEO** (Losing Electrons is Oxidation)

Reduction = gain of electrons—**GER** (Gaining Electrons is Reduction)

Example 1

Consider the following redox reaction.

$Mg + S \rightarrow MgS$

$Mg^0 + S^0 \rightarrow Mg^{2+}S^{2-}$ oxidation numbers assigned

In this reaction, Mg is the substance being oxidized because it is losing electrons. This should be familiar to you from the last lesson.

$Mg^0 \rightarrow Mg^{2+} + 2\ e^-$

On the other hand, S^0 is the substance being reduced because it is gaining electrons.

$S^0 + 2\ e^- \rightarrow S^{2-}$

In this case, Mg^0 is called the **oxidized substance**. If a substance gives up electrons readily, it is also said to be a **reducing agent** because it is allowing another substance to be reduced. In this reaction, Mg causes S to be reduced.

S is the **reduced substance**. S is also called the **oxidizing agent** because it causes Mg to be oxidized. If a substance gains electrons readily, it is said to be an **oxidizing agent**.

Example 2

Consider the following redox reaction.

$2\ Na + S \rightarrow Na_2S$

$2\ Na^0 + S^0 \rightarrow Na_2^+S^{2-}$

Na is oxidized from Na^0 to Na^+ because it has lost electrons to produce the Na^+ ion. Consequently, Na is the reducing agent because it causes S^0 to be reduced.

S is reduced from S^0 to S^{2-} because it has gained electrons to produce the S^{2-} ion. Consequently, S is the oxidizing agent because it causes Na^0 to be oxidized.

Oxidation Number Changes

In the previous lesson, oxidation numbers were described as useful tools when determining what has been oxidized and what has been reduced. Changes in oxidation numbers indicate that a redox reaction has taken place. The following website illustrates a redox reaction at the molecular level when a metal is immersed in an aqueous ionic solution: http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/redox/home.html. Watch the animation and see if you can predict what will occur just before the metal is placed in solution.

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

In a redox reaction, the oxidation number of one element will increase, whereas the oxidation number of another element will decrease. Not all aqueous reactions are redox reactions. For example, double-replacement reactions and acid-base reactions are **not** redox reactions. However, many single-replacement reactions, combination reactions, decomposition reactions, and combustion reactions **are** redox reactions.

> **Note**
>
> An *increase* in the oxidation number of an atom signifies a loss of electrons and, therefore, **oxidation**.
>
> A *decrease* in the oxidation number of an atom signifies a gain of electrons and, therefore, **reduction**.

Example 3

Ionic equation: $Cu_{(s)} + 2\ AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2\ Ag_{(s)}$

In this reaction, identify the following:

1. **What is reduced?**
   Ag^+ gains an electron and is, therefore, reduced.
2. **What is oxidized?**
   Cu^0 loses two electrons and is, therefore, oxidized.
3. **What is the reducing agent?**
   Cu^0 is the electron donor and is, therefore, the reducing agent.
4. **What is the oxidizing agent?**
   Ag^+ is the electron acceptor and is, therefore, the oxidizing agent.

Net ionic equation:

Electrons lost (oxidation) $Cu^0 \rightarrow Cu^{2+} + 2e^{1-}$

$$Cu^0{}_{(s)} + 2Ag^+{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Ag^0{}_{(s)}$$

Electrons gained (reduction) $2Ag^+ + 2e^{1-} \rightarrow 2Ag^0$

A reducing agent causes the reduction of another species.

An oxidizing agent causes the oxidation of another species.

The substance being reduced, Ag^+, is the oxidizing agent (also called an **electron acceptor**).

The substance being oxidized, Cu^0, is the reducing agent (also called an **electron donor**).

Example 4

If you have difficulty identifying when oxidation and reduction have taken place, use the steps outlined in this sample problem to help you.

Sample Problem

For the reaction, $Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$

1. State whether it is a redox or non-redox reaction.
2. If it is a redox reaction, identify the oxidized substance, the reduced substance, the oxidizing agent, and the reducing agent.

Solution

Step 1: Assign oxidation numbers to each substance based on the rules for assigning oxidation numbers.

$$\overset{0}{Zn}_{(s)} + \overset{+2}{Cu^{2+}}{}_{(aq)} \rightarrow \overset{+2}{Zn^{2+}}{}_{(aq)} + \overset{0}{Cu}_{(s)}$$

Step 2: Check to see which reactant is losing electrons. This will be the oxidized substance.

$$\overset{0}{Zn}_{(s)} + \overset{+2}{Cu^{2+}}{}_{(aq)} \rightarrow \overset{+2}{Zn^{2+}}{}_{(aq)} + \overset{0}{Cu}_{(s)}$$

Zn is losing 2 electrons to form Zn^{2+}. Therefore, Zn is oxidized. Zn is also the reducing agent because it supplies electrons to the reactant getting reduced.

Step 3: Check to see which reactant is gaining electrons. This will be the reduced substance.

$$\overset{0}{Zn_{(s)}} + \overset{+2}{Cu^{2+}_{(aq)}} \rightarrow \overset{+2}{Zn^{2+}_{(aq)}} + \overset{0}{Cu_{(s)}}$$

Cu^{2+} is gaining 2 electrons to form Cu. Therefore, Cu^{2+} is reduced. Cu^{2+} also is the oxidizing agent as it takes away electrons from the reactant being oxidized.

Step 4: Check to see if a reduction and an oxidation occur. If both processes occur, then it is a redox reaction.

Steps 2 and 3 confirm that this is a redox reaction.

Learning Activity 1.9: Redox Reactions

1. In each of the following reactions:
   a. Assign oxidation numbers.
   b. Identify the element oxidized, the element reduced, the oxidizing agent, and the reducing agent.

      A. $4\ Al_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Al_2O_{3(s)}$

      B. $2\ HNO_{3(aq)} + 6\ HI_{(aq)} \rightarrow 2\ NO_{(g)} + 3\ I_{2(s)} + 4\ H_2O_{(l)}$

      C. $4\ Al_{(s)} + 3\ MnO_{2(s)} \rightarrow 2\ Al_2O_{3(s)} + 3\ Mn_{(s)}$

      D. $2\ KClO_{3(s)} \rightarrow 2\ O_{2(g)} + 3\ KCl_{(aq)}$

   a. Assign oxidation numbers to each element using the blanks provided.

      A. $4\ \overset{\_\_}{Al}_{(s)} + 3\ \overset{\_\_}{O}_{2(g)} \rightarrow 2\ \overset{\_\_}{Al}_2\overset{\_\_}{O}_{3(s)}$

      B. $2\ \overset{\_\_}{H}\,\overset{\_\_}{N}\,\overset{\_\_}{O}_{3(aq)} + 6\ \overset{\_\_}{H}\,\overset{\_\_}{I}_{(aq)} \rightarrow 2\ \overset{\_\_}{N}\overset{\_\_}{O}_{(g)} + 3\ \overset{\_\_}{I}_{2(s)} + 4\ \overset{\_\_}{H}_2\overset{\_\_}{O}_{(l)}$

continued

Learning Activity 1.9: Redox Reactions (continued)

C. $4\ \overset{\_\_}{\text{Al}}_{(s)} + 3\ \overset{\_\_}{\text{Mn}}\overset{\_\_}{\text{O}}_{2(s)} \rightarrow 2\ \overset{\_\_}{\text{Al}}_2\overset{\_\_}{\text{O}}_{3(s)} + 3\ \overset{\_\_}{\text{Mn}}_{(s)}$

D. $2\ \overset{\_\_}{\text{K}}\ \overset{\_\_}{\text{Cl}}\ \overset{\_\_}{\text{O}}_{3(s)} \rightarrow 2\ \overset{\_\_}{\text{O}}_{2(g)} + 3\ \overset{\_\_}{\text{K}}\ \overset{\_\_}{\text{Cl}}_{(aq)}$

b. Identify the elements in the table below.

	The oxidized element is:	The reduced element is:	The oxidizing agent is:	The reducing agent is:
A.				
B.				
C.				
D.				

2. Consider the following two reactions. Are they both examples of redox reactions? Why or why not?
   - a. $\text{NaOH}_{(aq)} + \text{HCl}_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{NaCl}_{(aq)}$
   - b. $2\ \text{K}_{(s)} + 2\ \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_{2(g)} + 2\ \text{KOH}_{(aq)}$
3. Define oxidation and reduction in terms of change in oxidation number.
4. Complete this paragraph with the correct terms:

An oxidation is always accompanied by a _____________________ reaction.

The substance that does the oxidizing is called the ____________________

agent and is, therefore, ______________________. The substance that does

the reducing is called the _______________________ agent and is, therefore,

____________________________.

Check the answer key.

Lesson Summary

In this lesson, you learned how to identify the components of a redox reaction, including the oxidizing agent, reducing agent, oxidized substance, and reduced substance. You also learned some rules to help you assign oxidation numbers. In the next lesson, you will learn how to balance oxidation-reduction reactions in acidic and basic solutions.

Notes

Assignment 1.7: Interpreting Redox Reactions (10 marks)

1. Consider the following unbalanced equation:

   $HNO_{3(aq)} + HI_{(aq)} \rightarrow NO_{(g)} + 2\,H_2O_{(l)} + I_{2(aq)}$

   In this reaction, assign oxidation numbers to help you answer the following questions.

   a. What element is reduced? *(1 mark)*

   ____

   b. What element is oxidized? *(1 mark)*

   ____

   c. What is the reducing agent? *(1 mark)*

   ____

   d. What is the oxidizing agent? *(1 mark)*

   ____

2. Consider the following unbalanced equation:

   $2\,HBr_{(aq)} + Cl_{2(g)} \rightarrow 2\,HCl_{(aq)} + Br_{2(l)}$

   In this reaction, assign oxidation numbers to help you answer the following questions.

   a. What element is reduced? *(1 mark)*

   ____

   b. What element is oxidized? *(1 mark)*

continued

Assignment 1.7: Interpreting Redox Reactions (continued)

c. What is the reducing agent? *(1 mark)*

d. What is the oxidizing agent? *(1 mark)*

3. How does the change in oxidation number indicate oxidation or reduction? *(2 marks)*

LESSON 8: BALANCING REDOX REACTIONS (4 HOURS)

Lesson Focus

SLO C12-1-11: Balance oxidation-reduction reactions using redox methods.
Include: acidic and basic solutions

Lesson Introduction

In the previous lesson, you learned how to differentiate a redox reaction from a non-redox reaction. You learned how to identify the oxidizing agent, the reducing agent, the oxidized substance, and the reduced substance. In this lesson, you will practice balancing redox reactions in acidic and basic solutions.

Balancing Redox Reactions Using the Oxidation-Number Change Method

You might recall learning in Grade 10 Science that atoms are conserved in a chemical reaction. This means that the number of atoms in the reactants is equal to that of the products formed. You have already learned in this module that the number of electrons is also conserved, meaning the number lost is equal to the number gained. These changes can be identified by comparing oxidation numbers.

Balancing Electrons

The following example will illustrate why it is important to balance electrons lost and gained in redox equations. Consider the following reaction.

$$\mathrm{Cu} + \mathrm{Cl_2} \rightarrow \mathrm{Cu^+} + \mathrm{Cl^-}$$

If you balance this equation by inspection, you would balance the Cl and obtain the following:

$$\mathrm{Cu} + \mathrm{Cl_2} \rightarrow \mathrm{Cu^+} + 2\,\mathrm{Cl^-}$$

Now assign oxidation numbers and look at the electrons lost and gained.

Lose 1 e⁻

$$\overset{0}{Cu} + \overset{0}{Cl_2} \rightarrow \overset{+1}{Cu^+} + \overset{-1}{2Cl^-}$$

Gain 1e⁻ x 2 = 2e⁻

Notice that the atoms are balanced, but the electrons are not balanced. When you balance the electrons you get the following:

Lose 1e⁻ x 2 = 2e⁻

$$\overset{0}{2Cu} + \overset{0}{Cl_2} \rightarrow \overset{+1}{2Cu^+} + \overset{-1}{2Cl^-}$$

Gain 1e⁻ x 2 = 2e⁻

Notice that the charges are now balanced as well. The left side has no charge while the right side has 2 positives and 2 negatives. Both sides are now neutral.

In this lesson, you will learn how to balance redox reactions by comparing the increases and decreases in oxidation numbers. For this reason, the method is called the **oxidation-number change method**. This balancing method is based on the fact that the total number of electrons gained in reduction must equal the total number of electrons lost in oxidation (comparing the increases and decreases in oxidation numbers).

Example 1:

Follow the steps given to balance the chemical equation below.

$$P_{(s)} + HNO_{3(aq)} + H_2O_{(l)} \rightarrow NO_{(g)} + H_3PO_{4(aq)}$$

1. Assign oxidation numbers to all the atoms in the equation. Write the number above the appropriate atoms.

$$\overset{0}{P_{(s)}} + \overset{+1+5-2}{HNO_{3(aq)}} + \overset{+1\ -2}{H_2O_{(l)}} \rightarrow \overset{+2-2}{NO_{(g)}} + \overset{+1+5-2}{H_3PO_{4(aq)}}$$

2. Identify which atoms are oxidized and which are reduced. Use a line to connect the atoms that undergo oxidation and those that undergo reduction. Write the number of electrons lost/gained at the midpoint of each line.

5 electrons lost (oxidation)

$$\overset{0}{P}_{(s)} + \overset{+1+5-2}{HNO_3}_{(aq)} + \overset{+1\ -2}{H_2O}_{(l)} \rightarrow \overset{+2-2}{NO}_{(g)} + \overset{+1+5-2}{H_3PO_4}_{(aq)}$$

3 electrons gained (reduction)

3. Make the total increase in oxidation number equal to the total decrease in oxidation number by using appropriate coefficients.

3 x (+5 electrons) = +15

$$\overset{0}{P}_{(s)} + \overset{+1+5-2}{HNO_3}_{(aq)} + \overset{+1\ -2}{H_2O}_{(l)} \rightarrow \overset{+2-2}{NO}_{(g)} + \overset{+1+5-2}{H_3PO_4}_{(aq)}$$

5 x (–3 electrons) = –15

4. Place the coefficient 3 in front of the $P_{(s)}$ and H_3PO_4 and place the coefficient 5 in front of HNO_3 and NO.

$$3P_{(s)} + 5HNO_{3(aq)} + H_2O_{(l)} \rightarrow 5NO_{(g)} + 3H_3PO_{4(aq)}$$

5. Balance all other atoms as you normally would and do a final check to see if all atoms and charges are balanced. In this case, we need to add a coefficient of 2 in front of the water.

$$3P_{(s)} + 5HNO_{3(aq)} + 2H_2O_{(l)} \rightarrow 5NO_{(g)} + 3H_3PO_{4(aq)}$$

Learning Activity 1.10: Using the Oxidation-Number Change Method

Balance the following redox reactions using the oxidation-number change method.

1. $HNO_{3(aq)} + H_3AsO_{3(aq)} \rightarrow NO(g) + H_3AsO_{4(aq)} + H_2O_{(l)}$
2. $Cu_{(s)} + HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + NO_{(g)} + H_2O_{(l)}$

Check the answer key.

Balancing Redox Reactions: The Half-Reaction Method

We have used the oxidation-number method to balance redox reactions. When balancing a redox reaction in an acidic or basic medium, there are additional steps. Recall that, most of the time, an acidic substance contains H+, while a basic solution contains OH-. In these cases, the **half-reaction method** is generally more useful, and many people find it easier because there is no need to assign oxidation numbers. Whether acidic or basic, this method follows the same steps; however, there is one additional step at the end when dealing with a reaction that occurs in basic conditions.

Balancing Redox Reactions in Acidic Environments

Sample Problem:

Balance this redox reaction, which occurs in acidic conditions.

$$I^-_{(aq)} + KClO_{3(aq)} \rightarrow I_{2(aq)} + KCl_{(aq)}$$

Step 1: Write the net ionic equation for the reaction and remove any spectator ions. (We can remove state symbols for now.)

$$I^- + \cancel{K^+} + ClO_3^- \rightarrow I_2 + \cancel{K^+} + Cl^-$$

Step 2: Write the reduction half reaction and the oxidation half reaction separately.

> **Note**
>
> If you are unsure which reaction is which, you can assign oxidation numbers. Leave plenty of space between reactions and between molecules.

Oxidation: $I^- \rightarrow I_2$

Reduction: $ClO_3^- \rightarrow Cl^-$

Step 3: Balance all atoms **except any oxygen or hydrogen**.

Oxidation: $2I- \rightarrow I_2$

Reduction: $ClO_3^- \rightarrow Cl^-$

Step 4: Balance any **oxygen** by adding H_2O, and then any **hydrogen** by adding H+.

Oxidation: $2I^- \rightarrow I_2$

Reduction: $6H^+ + ClO_3^- \rightarrow Cl^- + 3H_2O$

Step 5: Total up the charge on both sides of each half-reaction, and **add electrons (e^-)** to balance the charges.

Individual charges:	(1- x 2) 2-		0		2-
Oxidation:	$2I^-$	$\rightarrow$	I_2	+	$2\ e^-$
Total charge:	2-	=		2-	

Individual charges:	6-		(6 x 1+) 6+		1-		1-		0
Reduction:	$6e^-$	+	$6H^+$	+	ClO_3^-	$\rightarrow$	Cl^-	+	$3H_2O$
Total charge:			1-			=		1-	

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 2 electrons and the reduction reaction has 6. We need to multiply the oxidation reaction by 3 in order to balance the electrons.

$$\begin{array}{lccccccccc}
\text{Oxidation:} & & & 3\text{ x} & [& 2I^- & \rightarrow & I_2 & + & 2\,e^-\ ] \\
\cline{6-10}
 & & & & & 6I^- & \rightarrow & 3I_2 & + & 6\,e^- \\
\text{Reduction:} & 6\,e^- & + & 6H^+ & + & ClO_3^- & \rightarrow & Cl^- & + & 3H_2O
\end{array}$$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons and any terms that appear on both sides.

$$\begin{array}{lcccccccccccc}
\text{Oxidation:} & & & & & 6I^- & \rightarrow & 3I_2 & + & \cancel{6\,e^-} & & & \\
\text{Reduction:} & \cancel{6\,e^-} & + & 6H^+ & + & ClO_3^- & \rightarrow & Cl^- & + & 3H_2O & & & \\
\hline
 & 6I^- & + & 6H^+ & + & ClO_3^- & \rightarrow & 3I_2 & + & Cl^- & + & 3H_2O &
\end{array}$$

You can do a final check by adding up charges on both sides to see if they balance.

$$\begin{array}{ccccccccccccc}
6I^- & + & 6H^+ & + & KClO_3 & \rightarrow & 3I_2 & + & KCl & + & 3H_2O \\
\overset{(6 \times 1-)}{6-} & + & 6+ & + & 0 & = & 0 & + & 0 & + & 0
\end{array}$$

Both sides add up to a charge of 0. This reaction is balanced.

Step 8: Add any spectator ions you removed in step 1 and balance atoms, if needed.

Balanced Redox Reaction:

$$6I^- + 6H^+ + KClO_3 \rightarrow 3I_2 + KCl + 3H_2O$$

Balancing Redox Reactions in Basic Environments

To balance basic solutions, follow the same steps as for acidic solutions. However, you need to add the same number of OH^- ions as H^+ ions to **both** sides of the equation. The H^+ ions form water.

$$H^+ + OH^- \rightarrow H_2O$$

Remember that, in this method, an acidic substance contains H^+, while a basic solution contains OH^-.

To summarize, the steps for balancing redox reactions in acidic environments are as follows:

1. Write the net ionic equation for the reaction and remove spectator ions.
2. Write the reduction half reaction and the oxidation half reaction separately.
3. Balance all atoms except oxygen or hydrogen.
4. Balance oxygen using H_2O, then any hydrogen using H^+.
5. Add up the charge on each side of each half-reaction and add electrons (e-) to balance the charges.
6. Multiply each half reaction by an appropriate number to balance the number of electrons between them.
7. Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons and any terms that appear on both sides.

For basic environments, add the following steps:

8. Add OH- ions to neutralize the H^+ ions, forming water. Add the same number of OH^- ions to the other side of the equation.
9. Cancel out H_2O molecules.
10. Add any spectator ions you removed in step 1.

Example:

Balance the following redox reaction, which occurs in a basic environment.

$$MnO_4^-{}_{(aq)} + C_2O_4^{2-}{}_{(aq)} \rightarrow MnO_{2(s)} + CO_3^{2-}{}_{(aq)}$$

Step 1: Write the net ionic equation for the reaction and remove any spectator ions. (In this case, this reaction is already written in net ionic form.)

$$MnO_4^- + C_2O_4^{2-} \rightarrow MnO_2 + CO_3^{2-}$$

Step 2: Write the reduction half reaction and the oxidation half reaction separately.

Note

If you are unsure which reaction is which you can assign oxidation numbers.

Checking oxidation numbers

$$\overset{+7\ \ -2}{MnO_4^-} + \overset{+3\ \ -2}{C_2O_4^{2-}} \rightarrow \overset{+4\ \ -2}{MnO_2} + \overset{+4\ \ -2}{C\,O_3^{2-}}$$

Oxidation: $C_2O_4^{2-} \rightarrow CO_3^{2-}$

Reduction: $MnO_4^- \rightarrow MnO_2$

Step 3: Balance all atoms **except any oxygen or hydrogen**.

Oxidation: $C_2O_4^{2-} \rightarrow 2CO_3^{2-}$

Reduction: $MnO_4- \rightarrow MnO_2$

Step 4: Balance any **oxygen first** by adding H_2O, and then any **hydrogen** by adding H^+.

Oxidation: $2H_2O + C_2O_4^{2-} \rightarrow 2CO_3^{2-} + 4H^+$

Reduction: $4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$

Step 5: Total up the charge on both sides of each half reaction, and **add electrons (e^-)** to balance the charges.

Individual charges: (2 x 2-) (4 x 1+)

Oxidation: $$\underline{\overset{0}{2H_2O} + \overset{2-}{C_2O_4^{2-}}} \rightarrow \underline{\overset{4-}{2CO_3^{2-}} + \overset{4+}{4H^+} + \overset{2-}{2e^-}}$$

Total charge: 2- = 2-

Individual charges: (4 x 1+)

Reduction: $$\underline{\overset{3-}{3e^-} + \overset{4+}{4H^+} + \overset{1-}{MnO_4^-}} \rightarrow \underline{\overset{0}{MnO_2} + \overset{0}{2H_2O}}$$

Total charge: 0 = 0

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 2 electrons and the reduction reaction has 3 electrons. We need to multiply the oxidation reaction by 3 and the reduction reaction by 2 in order to balance the electrons.

Oxidation: $3 \times [\ 2H_2O + C_2O_4^{2-} \rightarrow 2CO_3^{2-} + 4H^+ + 2e^-\ ]$

$6H_2O + 3C_2O_4^{2-} \rightarrow 6CO_3^{2-} + 12H^+ + 6e^-$

Reduction: $2 \times [\ 3e^- + 4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O\ ]$

$6e^- + 8H^+ + 2MnO_4^- \rightarrow 2MnO_2 + 4H_2O$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons and any terms that appear on both sides.

$$\begin{array}{llcl}
\text{Oxidation:} & & \overset{2H_2O}{\cancel{6H_2O}} + 3C_2O_4^{2-} \rightarrow 6CO_3^{2-} + \overset{4H^+}{\cancel{12H^+}} + \cancel{6e^-} \\
\text{Reduction:} & \cancel{6e^-} + \cancel{8H^+} & + 2MnO_4^{-} \rightarrow 2MnO_2 + \cancel{4H_2O} \\
\hline
 & & 2H_2O + 3C_2O_4^{2-} + 2MnO_4^{-} \rightarrow 6CO_3^{2-} + 2MnO_2 + 4H^+
\end{array}$$

Step 8: Add OH⁻ ions to neutralize the H⁺ ions. The same amount of OH⁻ is added to the other side.

In this case, we add 4OH⁻ to each side.

$$4OH^- + 2H_2O + 3C_2O_4^{2-} + 2MnO_4^- \rightarrow 6CO_3^{2-} + 2MnO_2 + \overbrace{\cancel{4H^+} + \cancel{4OH^-}}^{4H_2O}$$

Step 9: Cancel out H_2O.

$$\begin{array}{c}
4OH^- + \cancel{2H_2O} + 3C_2O_4^{2-} + 2MnO_4^- \rightarrow 6CO_3^{2-} + 2MnO_2 + \overset{4_2H_2O}{\cancel{4H_2O}} \\
\hline
4OH^- + 3C_2O_4^{2-} + 2MnO_4^- \rightarrow 6CO_3^{2-} + 2MnO_2 + 2H_2O
\end{array}$$

Do a final check by adding up the charges on both sides to see if they balance.

$$\begin{array}{ccccccccccccc}
4OH^- & + & 3C_2O_4^{2-} & + & 2MnO_4^- & \rightarrow & 6CO_3^{2-} & + & 2MnO_2 & + & 2H_2O \\
(4 \times 1-) & & (3 \times 2-) & & (2 \times 1-) & & (6 \times 2-) & & & & \\
4- & + & 6- & + & 2- & = & 12- & + & 0 & + & 0
\end{array}$$

Both sides add up to 12-. This reaction is balanced.

Step 10: Add any spectator ions you removed in step 1.

(We didn't have any in this case.)

Balanced Redox Reaction:

$$4OH^- + 3C_2O_4^{2-} + 2MnO_4^- \rightarrow 6CO_3^{2-} + 2MnO_2 + 2H_2O$$

Learning Activity 1.11: Acidic and Basic Environments

Balance the following reactions using the half-reaction method.

a) $Mn^{2+}_{(aq)} + BiO_3^{-}{}_{(aq)} \rightarrow MnO_4^{-}{}_{(aq)} + Bi^{2+}_{(aq)}$ (acidic conditions)

b) $N_2O_{(g)} + ClO^{-}_{(aq)} \rightarrow NO_2^{-}{}_{(aq)} + Cl^{-}_{(aq)}$ (basic conditions)

c) $NO_3^{-}{}_{(aq)} + Cl^{-}_{(aq)} \rightarrow NO_{(g)} + ClO^{-}_{(aq)}$ (acidic conditions)

d) $Cr^{3+}_{(aq)} + ClO^{-}_{(aq)} \rightarrow CrO_4^{2-}{}_{(aq)} + Cl^{-}_{(aq)}$ (basic conditions)

e) $I_{2(s)} + Na_2S_2O_{3(aq)} \rightarrow Na_2S_2O_{4(aq)} + NaI_{(aq)}$ (acidic conditions)

Check the answer key.

Lesson Summary

In this lesson, you learned how to balance redox reactions by comparing the increases and decreases in oxidation numbers (called the **oxidation-number change method**). This balancing method is based on the fact that the total number of electrons gained in reduction must equal the total number of electrons lost in oxidation. In the next lesson, you will conclude your study of redox reactions by investigating their practical applications.

Assignment 1.8: Balancing Redox Reactions (30 marks)

1. Balance the following redox reactions using the oxidation-number change method. Each reaction is worth 5 marks. You will earn marks for the following steps:
   - assigning correct oxidation numbers *(1 mark)*
   - showing electron loss *(1 mark)*
   - showing electron gain *(1 mark)*
   - assigning correct coefficients *(1 mark)*
   - simplifying terms that appear on both sides of the equation *(1 mark)*

   a. $SO_{2(g)} + HNO_{2(aq)} \rightarrow H_2SO_{4(aq)} + NO_{(g)}$

   b. $HNO_{3(aq)} + H_2S_{(aq)} \rightarrow NO(g) + S_{(s)} + H_2O_{(l)}$

continued

Assignment 1.8: Balancing Redox Reactions (continued)

2. Balance the following redox reactions using the half-reaction method. Each equation is worth 10 marks. You will earn marks for the following steps:
   - writing out the reduction and oxidation reactions separately *(2 marks)*
   - balancing the atoms in each half reaction *(2 marks)*
   - balancing the charge in each half reaction *(2 marks)*
   - balancing the electrons between the oxidation and reduction reactions *(1 mark)*
   - adding the reactions *(1 mark)*
   - simplifying terms (and adding OH- in basic reactions) *(1 mark)*
   - having a complete balanced reaction *(1 mark)*

   a. $Sn^{2+}{}_{(s)} + IO_3{}^{-}{}_{(aq)} \rightarrow Sn^{4+}{}_{(aq)} + I^{-}{}_{(aq)}$ (acidic solution)

   b. $Cl_{2(aq)} + S_2O_3{}^{2-}{}_{(aq)} \rightarrow SO_4{}^{2-}{}_{(aq)} + Cl^{-}{}_{(aq)}$ (basic solution)

LESSON 9: REDOX APPLICATIONS (3 HOURS)

Lesson Focus

SLO C12-1-12: Research practical applications of redox reactions.

Examples: rocket fuels, fireworks, household bleach, photography, metal recovery from ores, steel making, aluminum recycling, fuel cells, batteries, tarnish removal, fruit clocks, forensic blood detection using luminol, chemiluminescence/bioluminescence, electrolytic cleaning, electrodeposition, photochemical etching, antioxidants/preservatives...

Lesson Introduction

In the previous few lessons, you learned about redox reactions, including how to identify them, assigning oxidation numbers, and balancing methods. In this lesson, you will take a look at redox reactions that occur in the world around you.

Research a Topic

Over the remainder of this course, you will compile information that relates to one application of a redox reaction. You will need time and more study of chemistry to understand the full scope of the topic that you choose. The completed project will be handed in at the end of the module on electrochemistry (Module 6).

As you research your topic, focus on the following:

- the redox reaction taking place
- the effect of the process on the environment
- the energy consumption involved

Sources of information for your research could include the following:

- email communication to experts in the field
- interviews
- magazines
- journals
- newspapers
- websites

Redox Applications

The following information serves as an introduction to various applications of redox reactions. Hopefully, you will find one of particular interest that you can further investigate.

Rocket Fuels

Each solid rocket booster of the space shuttle, used during the first two minutes of the launch, contains 495 000 kg of an explosive mixture of ammonium perchlorate and aluminum.

$$NH_4ClO_{4(s)} + Al_{(s)} \rightarrow Al_2O_{3(g)} + HCl_{(g)} + N_{2(g)} + H_2O_{(g)}$$

Fireworks

The heat and thrust of fireworks' shells are produced by exothermic redox reactions. Fireworks are typically composed of an oxidizer such as potassium perchlorate, a fuel such as aluminum or magnesium, a binder, and some chemicals for the special effects of colour, sparks, and smoke. For example, green fireworks are made by adding a barium compound, and gold sparks are produced by adding iron filings or charcoal.

Household Bleach

Through the process of oxidation, unwanted colours (stains) are removed (oxidized) by the bleach. Colour is caused by the movement of electrons between different energy levels of the atoms of the material.

$OCl^-_{(aq)}$ + coloured stain molecule$_{(s)}$ → $Cl^-_{(aq)}$ + colourless oxidized stain molecule$_{(s)}$

Photography (non-digital)

There are three different redox reactions in black and white photography.

1. The film negative is an emulsion of silver bromide.
   $Ag^+_{(aq)} + Br^-_{(aq)} \rightarrow AgBr_{(s)}$
2. The film is processed and the remaining $Ag^+_{(aq)}$ is converted to free silver by a reducing agent. The unreacted AgBr is removed by an appropriate solution process. You now have the negative.
3. The negative is printed onto photographic paper.

Metal Recovery from Ores

Aluminum metal is obtained by the electrolysis of aluminum oxide (refined bauxite), using the Hall-Heroult process.

Cathode: $Al^{3+}_{(aq)} + 3e^{-} \rightarrow Al_{(s)}$

Anode: $2O_{2}{}^{-}{}_{(aq)} \rightarrow O_{2(g)} + 4e^{-}$

Net cell reaction: $4Al^{3+}_{(aq)} + 6O^{2-}_{(aq)} \rightarrow 4Al_{(s)} + 3O_{2(g)}$

This process uses huge amounts of electric energy. The recycling of aluminum is a lot more cost-effective than the processing of bauxite.

Copper, silver, gold, platinum, and palladium are the only transition metals that are unreactive enough to be found in nature uncombined with other elements.

Steelmaking

One aspect of steelmaking is the basic "oxygen process" used to purify iron (the most common method used). Scrap steel is mixed with molten iron in a blast furnace and then oxygen is introduced (injected) to oxidize the impurities.

Aluminum Recycling

All aluminum products can be recycled after use. Scrap is generally taken by road to the recycling plant where it is checked and sorted to determine composition and value. If the scrap is of unknown quality, the aluminum will first go through large magnets to remove any ferrous metal. Depending on the type of contamination present, some scrap must be processed further. Beverage cans, for example, must have their lacquer removed prior to re-melting.

Fuel Cells

The most common fuel cell is the hydrogen-oxygen fuel cell used in the space shuttle.

$$\begin{array}{ll} \text{oxidation:} & 2H_{2(g)} + 4OH^{-}_{(aq)} \rightarrow 2H_2O_{(l)} + 4e^{-} \\ \text{reduction:} & O_{2(g)} + 2H_2O_{(l)} + 4e^{-} \rightarrow 4OH^{-}_{(aq)} \\ \hline \text{overall:} & 2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)} \end{array}$$

Batteries

The flow of electrons in a battery is possible because zinc is oxidized in the battery and MnO_2 is reduced. The chemical reactions that occur are:

oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

reduction: $2MnO_{2(s)} + 2NH_4^+{}_{(aq)} + 2e^- \rightarrow Mn_2O_{3(s)} + 2NH_{3(aq)} + H_2O_{(l)}$

Tarnish Removal

Silver tarnish is Ag_2S and is formed by a redox reaction involving environmental sulphides. To remove the tarnish, aluminum reacts in the following way:

$$3Ag_2S_{(s)} + 2Al_{(s)} \rightarrow Al_2S_{3(s)} + 6Ag_{(s)}$$

Fruit Clocks

Inserting two electrodes of differing metals into a piece of fruit (such as a lemon) and connecting them with wires will cause an electric current will flow to a basic liquid crystal display clock:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Pacemaker Cells

Engineered in the 1940s by a Canadian, Dr. John Hopps, the pacemaker sends electrical impulses to the heart muscle in order to correct heartbeat irregularities. The pacemaker obtains its energy from a battery that lasts seven years.

Antioxidants/Preservatives

Oxidation can cause the decay of food and other organic material (e.g., human skin). Antioxidants help reduce decay of some essential amino acids and the loss of some vitamins. Vitamin C, vitamin E, BHT (butylated hydroxytoluene), BHA (butylated hydroxyanisole), sulphites, and sulphur dioxide are examples of antioxidants that react more readily with oxygen than food does. This keeps the food from spoiling.

Corrosion Prevention

Paint or another protective coating is one way to protect steel structures from corrosion. Sacrificial anodes of magnesium or other active metals are also used to prevent corrosion.

Forensic Blood Detection

To perform a luminol test, criminologists simply spray a luminol mixture wherever they think there might be blood. If hemoglobin and the luminol mixture come in contact, the iron in the hemoglobin accelerates a reaction between the hydrogen peroxide and the luminol. In this oxidation reaction, the luminol loses nitrogen and hydrogen atoms while it gains oxygen atoms, resulting in a compound called 3-aminophthalate. The reaction leaves the 3-aminophthalate in an energized state—the electrons in the oxygen atoms are boosted to higher orbitals. The electrons quickly fall back to a lower energy level, emitting the extra energy as a light photon. With iron accelerating the process, the light is bright enough to see in a dark room.

Chemiluminescence/Bioluminescence

Most chemiluminescence methods involve only a few chemical components to actually generate light. Luminol chemiluminescence and peroxyoxalate chemiluminescence are both used in bioanalytical methods. In each system, a "fuel" is chemically oxidized to produce an excited state product. In many luminol methods, it is this excited product that emits the light for the signal. In peroxyoxalate chemiluminescence, the initial excited state product does not emit light at all. Instead, it reacts with another compound, often a compound also viable as a fluorescent dye, and it is this fluorophore that becomes excited and emits light.

Bioluminescence is light produced by a chemical reaction within an organism. At least two chemicals are required. The one that produces the light is generically called a **luciferin**, and the one that drives or catalyzes the reaction is called a **luciferase**.

Electrolytic Cleaning (e.g., the Titanic)

Coatings of salts from seawater on metal objects are removed by an electrochemical process. A voltaic cell is set up with a cathode that is the object itself and a stainless steel anode in a basic solution. Chloride ions are removed when the electric current is turned on.

In another process, bacteria convert sulphate ions to hydrogen sulphide gas and cause silver coins and bars to become coated with silver sulphide after long periods of time at the bottom of the ocean. In an electrolytic cell, the silver in silver sulphide can be reduced to silver metal and reclaimed.

Electrodeposition (e.g., making CDs)

The process used in electroplating is called electrodeposition. The item to be coated is placed into a container containing a solution of one or more metal salts. The item is then connected to an electrical circuit, forming the cathode (negative) of the circuit, while an electrode (typically of the same metal to be plated) forms the anode (positive). When an electrical current is passed through the circuit, metal ions in the solution are attracted to the item. The result is a layer of metal on the item; however, considerable skill is required to assure an evenly-coated finished product. This process is analogous to a galvanic cell acting in reverse.

Photochemical Etching

In this process, ultraviolet light is used to transfer a pattern onto a piece of metal. Chemicals are then applied to remove certain areas in the pattern, creating an intricate design on the metal.

Lesson Summary

In this lesson, you chose a specific example of a redox reaction and investigated its application(s) in the world around you. In the coming weeks, you will complete your research and write the report, using the marking rubric provided at the end of this module as a guide. The completed project will be handed in at the end of the course. Congratulations, you have completed Module 1!

Assignment 1.9: Researching Redox Reactions (24 marks)

Remember that your completed project will be submitted at the end of Module 6 (at the end of the course). Begin your research now and continue over the coming modules. Here are some suggestions to help you get started.

Step 1: Identify the redox application you have chosen.

Step 2: Research the redox reaction taking place, including the oxidizing agent, the reducing agent, the oxidized substance, and the reduced substance.

Step 3: Suggest possible effects the process might have on the environment, both positive and negative.

continued

Assignment 1.9: Researching Redox Reactions (continued)

Step 4: Comment on the energy consumption involved in this redox reaction.

Step 5: Describe the redox application you have chosen (e.g., where it is used).

On the following two pages, you will find the marking rubric for this research project. Read it over and use it as a guide to help you meet the goals of this project. You can even remove it and carry it through to the following modules as you continue your research.

Written Research Report Marking Rubric					
0 mark	**1 mark**	**2 marks**	**3 marks**	**4 marks**	**5 marks**
General Requirements					
The report does not fulfill any of the five requirements.	The report fulfills one of the five requirements.	The report fulfills two of the five requirements.	The report fulfills three of the five requirements.	The report fulfills four of the five requirements.	■ The report includes a title page. ■ The report is word-processed. ■ The report is double-spaced. ■ The report includes an alphabetized bibliography. ■ At least five sources are cited in the bibliography.
Content/Information					
The report does not fulfill any of the five requirements.	The report fulfills one of the five requirements.	The report fulfills two of the five requirements.	The report fulfills three of the five requirements.	The report fulfills four of the five requirements.	■ The information presented is clear, accurate, and concise. ■ All relevant information is completely presented and easily understood. ■ The report includes all redox reactions being utilized by the application. ■ The report identifies all substances being oxidized and reduced, as well as any oxidizing and reducing agents. ■ The report describes how the relevant redox reactions are used by humans/nature.

continued

<table>
<tr><th colspan="5">Written Research Report Marking Rubric (continued)</th></tr>
<tr><th>0 mark</th><th>1 mark</th><th>2 marks</th><th>3 marks</th><th>4 marks</th></tr>
<tr><td colspan="5">Organization</td></tr>
<tr><td>The report does not fulfill any of the four requirements.</td><td>The report fulfills one of the four requirements.</td><td>The report fulfills two of the four requirements.</td><td>The report fulfills three of the four requirements.</td><td>■ The clear and concise introduction includes a description of the topic of interest.
■ The body follows a logical pattern.
■ There are smooth transitions between paragraphs.
■ The conclusion includes a brief summary of the topic’s connection to redox chemistry.</td></tr>
<tr><td colspan="5">Mechanics (Grammar, Spelling, and Punctuation)</td></tr>
<tr><td>Frequent errors demonstrate no sign of proofreading.</td><td>Five or six errors found.</td><td>Three or four errors found.</td><td>One or two errors found.</td><td>No errors found. Great job!</td></tr>
<tr><th>0 mark</th><th>1 mark</th><th>2 marks</th><th colspan="2">3 marks</th></tr>
<tr><td colspan="5">Visual Aids (Pictures, Charts, and Diagrams)</td></tr>
<tr><td>The visual aids do not fulfill any of the three requirements.</td><td>The visual aids fulfill one of the three requirements.</td><td>The visual aids fulfill two of the three requirements.</td><td colspan="2">■ Time and effort were obviously put forth in the preparation of the visual aids.
■ The visual aids used in the report are neat and organized.
■ The visual aids help make the application more easily understood.</td></tr>
<tr><th>0 mark</th><th>1 mark</th><th>2 marks</th><th colspan="2">3 marks</th></tr>
<tr><td colspan="5">Creativity</td></tr>
<tr><td>The report shows no creativity.</td><td>The report shows little creativity.</td><td>The report shows some creativity.</td><td colspan="2">The creativity in the report is “above and beyond.”</td></tr>
</table>

MODULE 1 SUMMARY

Congratulations, you have finished the first module of this course. You have learned that water is essential for life on Earth. More than two-thirds of Earth’s surface is water. By weight, the cells in our bodies contain between 65 and 90 per cent water. Many of the reactions of life occur in water; consequently, we say they are aqueous. In this module, you also studied three types of aqueous reactions: precipitation reactions, neutralization reactions, and oxidation-reduction reactions.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 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:

☐ Cover Sheet for Module 1 (found at the end of the course Introduction)
☐ Assignment 1.1: Solubility and Precipitation
☐ Assignment 1.2: Predicting Precipitate Formation
☐ Assignment 1.3: Writing Balanced Neutralization Reactions
☐ Assignment 1.4: Stoichiometry of Neutralization
☐ Assignment 1.5: Neutralization Calculations
☐ Assignment 1.6: Oxidation Numbers
☐ Assignment 1.7: Interpreting Redox Reactions
☐ Assignment 1.8: Balancing Redox Reactions

Remember that Assignment 1.9: Researching Redox Reactions will only be due at the end of the course, even though you began the work in this module!

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

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Notes

GRADE 12 CHEMISTRY (40S)

Module 1
Reactions in Aqueous Solutions

Learning Activity Answer Keys

MODULE 1:
REACTIONS IN AQUEOUS SOLUTIONS

Learning Activity 1.1: Introduction to Solutions

Answer each of the questions on solubility and precipitation in the space provided.

1. List two substances that are purchased as solids but are then made into solutions before use.

   Examples include, drink crystals, concentrated juices, soups and sauces, dishwashing and laundry detergents, bath bubbles, and others.

2. List some examples of solutions that you use in everyday life.

   Examples include, all the examples listed as answers in Question 1, as well as shampoo, milk, engine coolant, coffee, tea, vinegar, and others.

3. Explain why you think clothes might be easier to clean in hot water than in cold water.

   The solubility of laundry soap (being either solid or liquid in nature) increases as temperature increases. This means that the particles of the soap dissociate into anions and cations at an increased rate, and dissolve faster than in cold water.

4. Complete the dissociation equations for each of these ionic compounds.

   a. $NaBr_{(s)} \rightarrow \mathbf{Na^+_{(aq)} + Br^-_{(aq)}}$
   b. $HCl_{(aq)} \rightarrow \mathbf{H^+_{(aq)} + Cl^-_{(aq)}}$
   c. $Fe(OH)_{3(s)} \rightarrow \mathbf{Fe^{3+}_{(aq)} + 3\ OH^-_{(aq)}}$
   d. $Na_2(SO_4)_{(aq)} \rightarrow \mathbf{2Na^+_{(aq)} + SO_4{}^{2-}_{(aq)}}$

Learning Activity 1.2: Using Solubility Rules

Use the Table of Solubility Rules to predict which of the following substances have low solubility and which are soluble.

1. KBr **soluble**
2. $NaNO_3$ **soluble**
3. $BaSO_4$ **low solubility**
4. $FeCl_2$ **soluble**
5. tin (II) sulphate **soluble**
6. lithium carbonate **soluble**

7. silver iodide **low solubility**
8. ammonium bromide **soluble**
9. copper (II) chloride **solubility**
10. calcium sulphide **soluble**

Learning Activity 1.3: Predicting Precipitation

Write the complete set of equations (balanced chemical, total ionic, and net ionic) for each of the following reactions. Then, determine whether or not a precipitate will form in each case.

1. $MgCl_{2(aq)} + Na_3PO_{4(aq)}$
   - **Balanced chemical equation**
     $3\ MgCl_{2(aq)} + 2\ Na_3PO_{4(aq)} \rightarrow Mg_3(PO_4)_{2\ (s)} + 6\ NaCl_{(aq)}$
   - **From the solubility chart, you can also see that magnesium will have a low solubility with phosphate ions. Magnesium phosphate will likely form a precipitate.**
   - **After mixing, the two solutions will dissociate as follows:**
     $MgCl_{2(aq)} \rightarrow Mg^{2+}{}_{(aq)} + 2\ Cl^{-}{}_{(aq)}$
     $Na_3PO_{4(aq)} \rightarrow 3Na^{+}{}_{(aq)} + PO_4{}^{3-}{}_{(aq)}$

     The total ionic equation would be as follows:
     $3\ Mg^{2+}{}_{(aq)} + 6\ Cl^{-}{}_{(aq)} + 6\ Na^{+}{}_{(aq)} + 2PO_4{}^{3-}{}_{(aq)} \rightarrow Mg_3(PO_4)_{2(s)}$
     $+\ 6\ Cl^{-}{}_{(aq)} + 6Na^{+}{}_{(aq)}$
   - **Since Cl^{-} and Na^{+} appear on both sides of the equation, but are not involved in the reaction, they are called spectator ions. Spectator ions can be cancelled from the equation, leaving the following net ionic equation:**
     $3\ Mg^{2+}{}_{(aq)} + 2\ PO_4{}^{3-}{}_{(aq)} \rightarrow Mg_3(PO_4)_{2(s)}$

2. $Al_2(SO_4)_{3(aq)}$ and $Ba(NO_3)_{2(aq)}$
   - **Balanced chemical equation**
     $Al_2(SO_4)_{3(aq)} + 3\ Ba(NO_3)_{2(aq)} \rightarrow 2\ Al(NO_3)_{3(aq)} + 3\ BaSO_{4(s)}$
   - **From the solubility chart, you can also see that barium ions will have a low solubility with sulfate ions. Barium sulfate will likely form a precipitate.**

- **After mixing, the two solutions will dissociate as follows:**
  $Al_2(SO_4)_{3(aq)} \rightarrow 2\ Al^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)}$
  $Ba(NO_3)_{2(aq)} \rightarrow Ba^{2+}{}_{(aq)} + 2\ NO_3{}^{-}{}_{(aq)}$

  The total ionic equation would be as follows:
  $2\ Al^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)} + 3\ Ba^{2+}{}_{(aq)} + 6\ NO_3{}^{-}{}_{(aq)} \rightarrow 3\ BaSO_{4(s)} + 6\ NO_3{}^{-}{}_{(aq)} + 2\ Al^{3+}{}_{(aq)}$
- **The net ionic equation is as follows:**
  $SO_4{}^{2-}{}_{(aq)} + Ba^{2+}{}_{(aq)} \rightarrow BaSO_{4(s)}$
  (Remember to reduce coefficients to the lowest term when possible.)

3. $(NH_4)_2S_{(aq)}$ and $Co(NO_3)_{2(aq)}$
   - **Balanced chemical equation**
     $(NH_4)_2S_{(aq)}$ and $Co(NO_3)_{2(aq)} \rightarrow CoS_{(s)} + 2\ NH_4NO_{3(aq)}$
   - **From the solubility chart, you can also see that the cobalt ions will have a low solubility with the sulfide ions. Cobalt sulfide will likely form a precipitate.**
   - **The total ionic equation would be as follows:**
     $Co^{2+}{}_{(aq)} + 2\ NO_3{}^{-}{}_{(aq)} + 2\ NH_4{}^{+}{}_{(aq)} + S^{2-}{}_{(aq)} \rightarrow CoS_{(s)} + 2\ NO_3{}^{-}{}_{(aq)} + 2\ NH_4{}^{+}{}_{(aq)}$
   - **Spectator ions can be cancelled from the equation leaving the following net ionic equation:**
     $Co^{2+}{}_{(aq)} + S^{2-}{}_{(aq)} \rightarrow CoS_{(s)}$

Learning Activity 1.4: Neutralization Reactions

1. Complete and balance the following neutralization reactions. Be sure to include state symbols.

   a. $HBr_{(aq)} + NaOH_{(aq)} \rightarrow$ **$NaBr_{(aq)} + H_2O_{(l)}$**

   b. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow$ **$NaCl_{(aq)} + H_2O_{(l)}$**

   c. $H_2SO_{4(aq)} + KOH_{(aq)} \rightarrow$
      $H_2SO_{4(aq)} + 2\ KOH_{(aq)} \rightarrow K_2SO_{4(aq)} + 2\ H_2O_{(l)}$

   d. $NaOH_{(aq)} + HI_{(aq)} \rightarrow$ **$NaI_{(aq)} + H_2O_{(l)}$**

   e. $Ca(OH)_{2(aq)} + HCl_{(aq)} \rightarrow$
      $Ca(OH)_{2(aq)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + 2\ H_2O_{(l)}$

   f. $HNO_{3(aq)} + KOH_{(aq)} \rightarrow$ **$KNO_{3(aq)} + H_2O_{(l)}$**

2. What is a **strong base**? Give an example.

   A strong base completely dissociates into ions. For example, if there were 100 molecules of NaOH dissolved into water, 100 Na+ ions and 100 OH^- ions would be produced. Some examples of strong bases include sodium hydroxide (NaOH), potassium hydroxide (KOH), cesium hydroxide (CsOH), calcium hydroxide ($Ca(OH)_2$), and barium hydroxide ($Ba(OH)_2$).

3. What is a **salt**? Give an example.

   A salt is any ionic, crystalline compound that can be produced from the neutralization of a strong acid and a strong base. Examples include $CaCl_2$, NaCl, NaI, and NaBr.

Learning Activity 1.5: Titration

1. As you watch the animation, pause at 4:00 and 4:18 for the data given at the end of the titration and record the pH on the table below.

   By titration: Equivalence: 26.50 mL

Volume of NaOH (mL)	pH
0.00	**1.49**
5.00	**1.47**
10.00	**1.60**
15.00	**1.71**
20.00	**1.91**
21.00	**1.99**
22.00	**2.03**
23.00	**2.16**
24.00	**2.31**
25.00	**2.54**
26.00	**3.70**
27.00	**10.92**
28.00	**11.34**
29.00	**11.35**
30.00	**11.41**
35.00	**11.69**
40.00	**11.78**
45.00	**11.89**
50.00	**11.93**

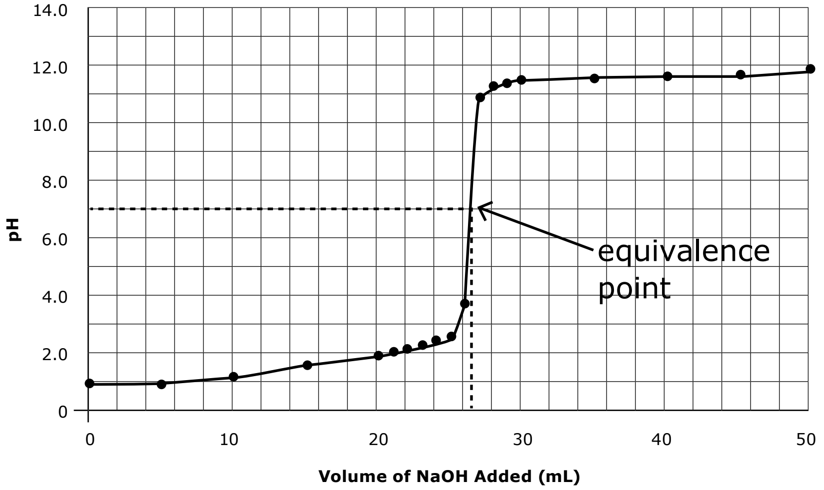
2. Write a balanced chemical equation for the reaction between HCl and NaOH.

   $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

3. Write the net ionic equation for the reaction between HCl and NaOH.

   $H^+_{(aq)} + OH_{(aq)} \rightarrow H_2O_{(l)}$

4. Graph the data. The pH will be on the vertical (*y*) axis and the volume of NaOH will be on the horizontal (*x*) axis. Be sure to label your axes and provide an appropriate title for your graph. Label the equivalence point on the graph.



pH versus Volume of NaOH

5. What volume of base was required for neutralization to occur? Explain how you know this.

   Neutralization occurred between 26 mL and 27 mL. This is where the pH drastically went from acidic to basic, passing over the neutral point. According to the pH measured, the solution composition changed suddenly from excess acid to excess base.

6. Calculate the number of moles of NaOH required to neutralize the sample of HCl.

   $NaOH = \frac{0.20 \text{ mol}}{L} \times 0.02650 \text{ NaOH} = 0.0053 \text{ mol}$

7. Calculate the number of moles present in 50.0 mL of HCl in the laboratory.

   $HCl = \frac{0.1068 \text{ mol}}{L} \times 0.0500 \text{ L} = 0.0053 \text{ mol}$

8. Using the balanced molecular equation, determine the ratio of moles between the sodium hydroxide and the hydrochloric acid.

   $\mathbf{HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}}$

   Coefficient NaOH/coefficient HCl = 1:1 = 1

9. Using the number of moles obtained in Questions 5 and 6, determine the experimental ratio of moles between the NaOH and the HCl.

   $\frac{\text{moles NaOH}}{\text{moles HCl}} = \frac{0.0053}{0.0053} = 1$

Learning Activity 1.6: Concentration Calculations

Use the problem-solving method outlined in this lesson to perform the following calculations. Show your work and make sure your answers include the correct units and significant digits.

1. Calculate the concentration of KOH if 60. mL of this base is neutralized by 25. mL of a 0.50 mol/L solution of HCl.
   a. Write the **balanced chemical equation** for the neutralization reaction.

   $\mathbf{HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}}$

	$HCl_{(aq)}$	$KOH_{(aq)}$
c (mol/L)	**0.50 mol/L**	**0.21 mol/L**
v (L)	**0.025 L**	**0.060 L**
n (mol)	**0.012 mol**	**0.012 mol**

   b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).

   $\mathbf{\frac{0.50 \text{ mol HCl}}{\text{L}} \times 0.025 \text{ L} = 0.025 \text{ mol HCl}}$

c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).

$$\frac{\textbf{mol HCl}}{\textbf{coefficient HCl}} = \frac{\textbf{mol KOH}}{\textbf{coefficient KOH}}$$

$$\frac{\textbf{0.0125 mol HCl}}{\textbf{1 mol HCl}} = \frac{\textbf{mol KOH}}{\textbf{1 mol KOH}}$$

0.0125 mol KOH = mol KOH

d. Calculate the **required concentration** of the base.

$$\textbf{C} = \frac{\textbf{moles}}{\textbf{volume}} = \frac{\textbf{0.0125 mol KOH}}{\textbf{0.06 L}} = \textbf{0.21 mol/L}$$

2. Calculate the concentration of H_2SO_4 if 60.0 mL of this acid is neutralized by 10 mL of a 2.0 mol/L solution of NaOH.

a. Write the **balanced chemical equation** for the neutralization reaction.

$$H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2\,H_2O_{(l)}$$

	$H_2SO_{4(aq)}$	2 $NaOH_{(aq)}$
c (mol/L)	**0.17 mol/L**	**2.0 mol/L**
v (L)	**0.060 L**	**0.010 L**
n (mol)	**0.01 mol**	**0.02 mol**

b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).

$$\frac{\textbf{2.0 mol NaOH}}{\textbf{L}} \times \textbf{0.010 L} = \textbf{0.020 mol NaOH}$$

c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).

$$\frac{\textbf{mol } H_2SO_4}{\textbf{coefficient } H_2SO_4} = \frac{\textbf{mol NaOH}}{\textbf{coefficient NaOH}}$$

$$\frac{\textbf{mol } H_2SO_4}{\textbf{1 mol } H_2SO_4} = \frac{\textbf{0.02 mol NaOH}}{\textbf{2 mol NaOH}}$$

0.01 mol H_2SO_4 = mol H_2SO_4

d. Calculate the **required concentration** of the acid.

$$\mathbf{C} = \frac{\textbf{moles}}{\textbf{volume}} = \frac{\textbf{0.01 mol}}{\textbf{0.06 L}} = \textbf{0.17 mol/L} = \textbf{0.2 mol/L}$$

3. How many mL of 0.50 mol/L KOH will exactly neutralize 15.0 mL of a 2.0 mol/L solution of H_2SO_4?

a. Write the **balanced chemical equation** for the neutralization reaction.

$$H_2SO_{4(aq)} + 2\,KOH_{(aq)} \rightarrow K_2SO_{4(aq)} + 2\,H_2O_{(l)}$$

	2KOH	H_2SO_4
c (mol/L)	**0.50 mol/L**	**2.0 mol/L**
v (L)	**0.12 L**	**0.015 L**
n (mol)	**0.060 mol**	**0.030 mol**

b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).

$$\frac{2.0 \text{ mol } H_2SO_4}{L} \times 0.015 \text{ L} = 0.030 \text{ mol } H_2SO_4$$

c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).

$$\frac{\text{mol } H_2SO_4}{\text{coefficient } H_2SO_4} = \frac{\text{mol KOH}}{\text{coefficient KOH}}$$

$$\frac{0.030 \text{ mol } H_2SO_4}{1 \text{ mol } H_2SO_4} = \frac{\text{mol KOH}}{2 \text{ mol KOH}}$$

0.060 mol KOH = mol KOH

d. Calculate the **required volume** of the base.

$$V = \frac{\text{moles}}{C} = \frac{0.060 \text{ mol KOH}}{0.50 \text{ mol/L}} = 0.12 \text{ mol/L}$$

4. How many mL of 1.0 mol/L $Ba(OH)_2$ solution must be added to 45.1 mL of a 0.20 mol/L HBr solution to make a neutral solution?

   a. Write the **balanced chemical equation** for the neutralization reaction.

   $$Ba(OH)_{2(aq)} + 2\,HBr_{(aq)} \rightarrow BaBr_{2(aq)} + 2\,H_2O_{(l)}$$

	$Ba(OH)_{2(aq)}$	$2\,HBr_{(aq)}$
c (mol/L)	**1.0 mol/L**	**0.20 mol/L**
v (L)	**0.0045 L**	**0.0451 L**
n (mol)	**0.0045 mol**	**0.009 mol**

   b. Use the concentration and volume of the known acid or base to **calculate the moles** of the **known substance** (acid or base).

   $$\frac{0.20 \text{ mol HBr}}{\text{L}} \times 0.0451 \text{ L} = 0.0090 \text{ mol HBr}$$

   c. Use the coefficients from the balanced equation to determine the **moles of the unknown substance** (acid or base).

   $$\frac{\text{mol HBr}}{\text{coefficient HBr}} = \frac{\text{mol Ba}\left(\text{OH}_2\right)}{\text{coefficient Ba}\left(\text{OH}_2\right)}$$

   $$\frac{0.0090 \text{ mol HBr}}{2 \text{ mol HBr}} = \frac{\text{mol Ba}\left(\text{OH}_2\right)}{1 \text{ mol Ba}\left(\text{OH}_2\right)}$$

   $$0.0045 \text{ mol Ba}\left(\text{OH}_2\right) = \text{mol Ba}\left(\text{OH}_2\right)$$

   d. Calculate the **required volume** of base.

   $$V = \frac{\text{moles}}{C} = \frac{0.0045 \text{ mol}}{1.0 \text{ mol/L}} = 0.0045 \text{ L} = 4.5 \text{ mL}$$

Learning Activity 1.7: Interpreting a Redox Reaction

1. If a neutral atom becomes positively charged, has it been oxidized or reduced? Why?

   If an ion becomes positively charged, it has been oxidized. For example, if an atom loses two electrons, it now has two MORE protons than electrons. When there are more protons than electrons, a charge imbalance results. This charge imbalance leads to the formation of a positively charged ion (a cation).

2. If an ion with a 1^- charge acquires a 2^- charge, has it been oxidized or reduced? Why?

   In this case, the ion has been reduced. The change in charge from 1- to 2- indicates that one electron was gained in the process. The ion, having an electron, now has one electron MORE than protons. This charge imbalance results in the formation of an even more negatively charged ion (an anion).

3. Consider this equation: $Mg^0 + Co^{2+} \rightarrow Co^0 + Mg^{2+}$
   a. Which substance is oxidized?

      magnesium (Mg)

   b. Which substance is reduced?

      cobalt Co^{2+} ion

4. Consider this equation: $Fe^{3+} + Cu^{1+} \rightarrow Fe^{2+} + Cu^{2+}$
   a. Which substance is oxidized?

      copper Cu^+ ion

   b. Which substance is reduced?

      iron Fe^{3+} ion

5. Why must oxidation and reduction occur together?

   Oxidation and reduction must occur together because the number of electrons lost must equal the number of electrons gained. This assures that the charge is conserved on both sides of the equation. Therefore, an oxidation-reduction reaction, or redox reaction, is a reaction in which electrons are transferred (lost or gained) between reactants.

6. Complete the following statements:

   In a balanced redox reaction, one substance is __**oxidized**__ when it loses electrons. In the same balanced redox reaction, the other substance is __**reduced**__ and it therefore gains electrons.

Learning Activity 1.8: Assigning Oxidation Numbers

1. What is the oxidation number of each element in the following?

   a. Fe **0**

   b. Ca^{2+} **+2**

   c. SO_2 **S = +4, O = –2**

   d. Na_2SO_4 **Na = +1, S = +6, O = –2**

   e. CO_3^{2-} **C = +4, O = –2**

   f. $MgCl_2$ **Mg = +2, Cl = –1**

2. In each of the following, determine the oxidation number of the element that is in **bold**.

   a. **Ag** **0**

   b. $\mathbf{Br}_2$ **0**

   c. $KN\mathbf{O}_3$ **–2**

   d. $K\mathbf{Cl}O_3$ **+5**

   e. $Pb\mathbf{S}O_4$ **+6**

   f. $\mathbf{K}_3PO_4$ **+1**

   g. $N\mathbf{O}_\mathbf{3}^-$ **–2**

   h. $\mathbf{N}H_4^+$ **–3**

   i. $\mathbf{Na}_2O_2$ **+1**

   j. $\mathbf{S}O_2$ **+4**

Learning Activity 1.9: Redox Reactions

1. In each of the following reactions:
   a. Assign oxidation numbers.
   b. Identify the element oxidized, the element reduced, the oxidizing agent, and the reducing agent.

   A. $4\,\text{Al}_{(s)} + 3\,\text{O}_{2(g)} \rightarrow 2\,\text{Al}_2\text{O}_{3(s)}$

   B. $2\,\text{HNO}_{3(aq)} + 6\,\text{HI}_{(aq)} \rightarrow 2\,\text{NO}_{(g)} + 3\,\text{I}_{2(s)} + 4\,\text{H}_2\text{O}_{(l)}$

   C. $4\,\text{Al}_{(s)} + 3\,\text{MnO}_{2(s)} \rightarrow 2\,\text{Al}_2\text{O}_{3(s)} + 3\,\text{Mn}_{(s)}$

   D. $2\,\text{KClO}_{3(s)} \rightarrow 2\,\text{O}_{2(g)} + 3\,\text{KCl}_{(aq)}$

a. Assign oxidation numbers to each element using the blanks provided.

A. $4\,\overset{\mathbf{0}}{\text{Al}}_{(s)} + 3\,\overset{\mathbf{0}}{\text{O}}_{2(g)} \rightarrow 2\,\overset{\mathbf{+3}}{\text{Al}}_2\overset{\mathbf{-2}}{\text{O}}_{3(s)}$

B. $2\,\overset{\mathbf{+1}}{\text{H}}\,\overset{\mathbf{+5}}{\text{N}}\,\overset{\mathbf{-2}}{\text{O}}_{3(aq)} + 6\,\overset{\mathbf{+1}}{\text{H}}\,\overset{\mathbf{-1}}{\text{I}}_{(aq)} \rightarrow 2\,\overset{\mathbf{+2}}{\text{N}}\overset{\mathbf{-2}}{\text{O}}_{(g)} + 3\,\overset{\mathbf{0}}{\text{I}}_{2(s)} + 4\,\overset{\mathbf{+1}}{\text{H}}_2\overset{\mathbf{-2}}{\text{O}}_{(l)}$

C. $4\,\overset{\mathbf{0}}{\text{Al}}_{(s)} + 3\,\overset{\mathbf{+4}}{\text{Mn}}\overset{\mathbf{-2}}{\text{O}}_{2(s)} \rightarrow 2\,\overset{\mathbf{+3}}{\text{Al}}_2\overset{\mathbf{-2}}{\text{O}}_{3(s)} + 3\,\overset{\mathbf{0}}{\text{Mn}}_{(s)}$

D. $2\,\overset{\mathbf{+1}}{\text{K}}\,\overset{\mathbf{+5}}{\text{Cl}}\,\overset{\mathbf{-2}}{\text{O}}_{3(s)} \rightarrow 2\,\overset{\mathbf{0}}{\text{O}}_{2(g)} + 3\,\overset{\mathbf{+1}}{\text{K}}\,\overset{\mathbf{-1}}{\text{Cl}}_{(aq)}$

b. Identify the elements in the table below.

	The oxidized element is:	The reduced element is:	The oxidizing agent is:	The reducing agent is:
A.	**Al**	**O_2**	**$\text{O}_{2(g)}$**	**$\text{Al}_{(s)}$**
B.	**I**	**N**	**$\text{HNO}_{3(aq)}$**	**$\text{HI}_{(aq)}$**
C.	**Al**	**Mn**	**$\text{MnO}_{2(s)}$**	**$\text{Al}_{(s)}$**
D.	**O**	**Cl**	**$\text{KClO}_{3(s)}$**	**$\text{KClO}_{3(s)}$**

2. Consider the following two reactions. Are they both examples of redox reactions? Why or why not?
   a. $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

      This is a non-redox reaction because there is no loss or gain of electrons. You can verify this by assigning oxidation numbers. The numbers do not change from reactant to product side.

   b. $2\ K_{(s)} + 2\ H_2O_{(l)} \rightarrow H_{2(g)} + 2\ KOH_{(aq)}$

      This is a redox reaction because the number of electrons lost does equal the number of electrons gained. You can verify this by assigning oxidation numbers. The numbers change from reactant to product side.

3. Define oxidation and reduction in terms of change in oxidation number.

   An increase in the oxidation number of an atom signifies oxidation, while a decrease in the oxidation number signifies reduction.

4. Complete this paragraph with the correct terms:

   An oxidation is always accompanied by a **reduction** reaction. The substance that does the oxidizing is called the **oxidizing** agent and is, therefore, **reduced**. The substance that does the reducing is called the **reducing** agent and is, therefore, **oxidized**.

Learning Activity 1.10: Using the Oxidation-Number Change Method

Balance the following redox reactions using the oxidation-number change method.

1. $HNO_{3(aq)} + H_3AsO_{3(aq)} \rightarrow NO_{(g)} + H_3AsO_{4(aq)} + H_2O_{(l)}$

   The N atoms change from +5 to +2, so they are reduced.

   The As atoms, change from +3 to +5, so they are oxidized.

   As: +3 to +5, Net Change = +2

   N: +5 to +2, Net Change = –3

   Six electrons would be lost by three arsenic atoms.

   2 N atoms would yield a net decrease of –6.

   Two nitrogen atoms would gain six electrons.

Add coefficients to the formulas which contain the elements whose oxidation number is changing.

$2\ HNO_{3(aq)} + 3\ H_3AsO_{3(aq)} \rightarrow NO_{(g)} + H_3AsO_{4(aq)} + H_2O_{(l)}$

Finally, balance the rest of the equation by inspection.

$2\ HNO_{3(aq)} + 3\ H_3AsO_{3(aq)} \rightarrow 2\ NO_{(g)} + 3\ H_3AsO_{4(aq)} + H_2O_{(l)}$

2. $Cu_{(s)} + HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + NO_{(g)} + H_2O_{(l)}$

The copper atoms are changing their oxidation number from 0 to +2.

The nitrogen atoms are changing from +5 to +2.

These changes indicate that this is a redox reaction.

Cu: 0 to +2, Net Change = +2

Some N: +5 to +2, Net Change = –3

(Nitrogen atoms change oxidation numbers: the 2 goes on the product side in NO due to the change in oxidation numbers.)

3 Cu atoms are needed (net change of +6) for every 2 nitrogen atoms that change (net change of –6). Place the 2 in front of the $NO_{(g)}$ on the product side (this is where the nitrogen atom ends up on the product side). The 3 for the copper atoms can be placed in front of the $Cu_{(s)}$.

$3\ Cu_{(s)} + 2\ HNO_{3(aq)} \rightarrow 3\ Cu(NO_3)_{2(aq)} + 2\ NO_{(g)} + H_2O_{(l)}$

Balance the rest of the atoms, being careful to keep the ratio of Cu to NO 3:2.

$3\ Cu_{(s)} + 8\ HNO_{3(aq)} \rightarrow 3\ Cu(NO_3)_{2(aq)} + 2\ NO_{(g)} + 4\ H_2O_{(l)}$

Learning Activity 1.11: Acidic and Basic Environments

Balance the following reactions using the half reaction method:

a) $Mn^{2+}{}_{(aq)} + BiO_{3^-(aq)} \rightarrow MnO_{4^-(aq)} + Bi^{2+}{}_{(aq)}$ (acidic conditions)

Step 1: Write the net ionic equation for the reaction and remove any spectator ions.

(In this case this reaction is already written in net ionic form).

$$Mn^{2+} + BiO_3^- \rightarrow MnO_4^- + Bi^{2+}$$

Step 2: Write the reduction half-reaction and the oxidation half-reaction separately.

(HINTS: If you are unsure which reaction is which you can assign oxidation numbers.)

Checking
oxidation numbers

$$\overset{+2}{Mn}{}^{2+} + \overset{+5}{Bi}\overset{-2}{O}{}_3^- \rightarrow \overset{+7}{Mn}\overset{-2}{O}{}_4^- + \overset{+2}{Bi}{}^{2+}$$

Oxidation: $BiO_3^- \rightarrow Bi^{2+}$

Reduction: $Mn^{2+} \rightarrow MnO_4^-$

Step 3: Balance all atoms **except any oxygen or hydrogen** (in this case they balance)

Oxidation: $BiO_3^- \rightarrow Bi^{2+}$

Reduction: $Mn^{2+} \rightarrow MnO_4^-$

Step 4: Balance any **oxygen first** by adding H_2O and then **hydrogen** by adding H^+

Oxidation: $6H^+ + BiO_3^- \rightarrow Bi^{2+} + 3H_2O$

Reduction: $4H_2O + Mn^{2+} \rightarrow MnO_4^- + 8H^+$

Step 5: Total up the charge on each side of each half-reaction, and **add electrons** (e^-) to balance the charges.

Individual Charges:
Oxidation:

$$\underline{\overset{3-}{3\,e^-} + \overset{6+}{6H^+} + \overset{1-}{BiO_3^-}} \rightarrow \underline{\overset{2+}{Bi^{2+}} + \overset{0}{3H_2O}}$$

Total charge: 2+ = 2+

Individual Charges:
Reduction:

$$\underline{\overset{0}{4H_2O} + \overset{2+}{Mn^{2+}}} \rightarrow \underline{\overset{1-}{MnO_4^-} + \overset{8+}{8H^+} + \overset{5-}{5\,e^-}}$$

Total charge: 2+ = 2+

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 3 electrons and the reduction reaction has 5. We need to multiply the oxidation reaction by 5 and the reduction reaction by 3 in order to balance the electrons.

Oxidation: $5 \times [\ 3e^- + 6H^+ + BiO_3^- \rightarrow Bi^{2+} + 3H_2O\ ]$

$15e^- + 30H^+ + 5BiO_3^- \rightarrow 5Bi^{2+} + 15H_2O$

Reduction: $3 \times [\ 4H_2O + Mn^{2+} \rightarrow MnO_4^- + 8H^+ + 5e^-\ ]$

$12H_2O + 3Mn^{2+} \rightarrow 3MnO_4^- + 24H^+ + 15e^-$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons, H_2O, and H+ that appear on both sides.

Oxidation: $\cancel{15e^-} + \overset{6H^+}{\cancel{30H^+}} + 5BiO_3^- \rightarrow 5Bi^{2+} + \overset{3H_2O}{\cancel{15H_2O}}$

Reduction: $\cancel{12H_2O} + 3Mn^{2+} \rightarrow 3MnO_4^- + \cancel{24H^+} + 15e^-$

$$6H^+ + 3Mn^{2+} + 5BiO_3^- \rightarrow 3MnO_4^- + 5Bi^{2+} + 3H_2O$$

Do a final check by adding up the charges on both sides to see if they balance:

$$\begin{array}{ccccccccccc} 6H^+ & + & 3Mn^{2+} & + & 5BiO_3^- & \rightarrow & 3MnO_4^- & + & 5Bi^{2+} & + & 3H_2O \\ (6 \times 1+) & & (3 \times 2+) & & (5 \times 1-) & & (3 \times 1-) & & (5 \times 2+) & & \\ 6+ & + & 6+ & + & 5- & = & 3- & + & 10+ & + & 0 \end{array}$$

Both sides add up to 7+. This reaction is balanced.

Step 8: Add any spectator ions you removed in step 1.

(We didn't have any in this case.)

Balanced Redox Reaction:

$$6H^+ + 3Mn^{2+} + 5BiO_3^- \rightarrow 3MnO_4^- + 5Bi^{2+} + 3H_2O$$

b) $N_2O_{(g)} + ClO^-_{(aq)} \rightarrow NO_2^-{}_{(aq)} + Cl^-_{(aq)}$ (basic conditions)

Step 1: Write the net ionic equation for the reaction and remove any spectator ions.

(In this case, this reaction is already written in net ionic form).

$$N_2O + ClO^- \rightarrow NO_2^- + Cl^-$$

Step 2: Write the reduction half-reaction and the oxidation half-reaction separately.

(HINT: If you are unsure which reaction is which you can assign oxidation numbers.)

Checking oxidation numbers	+1 -2		+1 -2		+3 -2		- 1
	N_2O	+	ClO^-	→	NO_2^-	+	Cl^-

Oxidation: $N_2O \rightarrow NO_2^-$

Reduction: ClO- $\rightarrow Cl^-$

Step 3: Balance all atoms **except any oxygen or hydrogen**.

In this case, balance the nitrogen:

Oxidation: $N_2O \rightarrow 2NO_2^-$

Reduction: $ClO^- \rightarrow Cl^-$

Step 4: Balance any **oxygen first** by adding H_2O and then **hydrogen** by adding H^+.

Oxidation: $3H_2O + N_2O \rightarrow 2NO_2^- + 6H^+$

Reduction: $2H^+ + ClO^- \rightarrow Cl^- + H_2O$

Step 5: Total up the charge on each side of each half-reaction, and **add electrons** (e^-) to balance the charges.

Individual Charges:					(2 x 2-)		(6 x 1+)		
	0		0		4-		6+		2-
Oxidation:	$3H_2O$	+	N_2O	→	$2NO_2^-$	+	$6H^+$	+	$2\,e^-$
Total charge:		0		=			0		

Individual Charges:			(2 x 1+)						
	2-		2+		1-		1-		0
Reduction:	$1\,e^-$	+	$2H^+$	+	ClO^-	→	Cl^-	+	H_2O
Total charge:			1-			=		1-	

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 2 electrons and the reduction reaction has 1. We do not need to multiply the oxidation reaction by any coefficient but the reduction reaction must be multiplied by 2 in order to balance the electrons.

Oxidation: $3H_2O + N_2O \rightarrow 2NO_2^- + 6H^+ + 2e^-$

Reduction: $2 \times [\ 1e^- + 2H^+ + ClO^- \rightarrow Cl^- + H_2O\ ]$

$$2e^- + 4H^+ + 2ClO^- \rightarrow 2Cl^- + 2H_2O$$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons and any terms that appear on both sides.

Oxidation: $\overset{H_2O}{\cancel{3H_2O}} + N_2O \rightarrow 2NO_2^- + \overset{2H^+}{\cancel{6H^+}} + \cancel{2e^-}$

Reduction: $\cancel{2e^-} + \cancel{4H^+} + 2ClO^- \rightarrow 2Cl^- + \cancel{2H_2O}$

$$H_2O + N_2O + 2ClO^- \rightarrow 2NO_2^- + 2Cl^- + 2H^+$$

Step 8: Because this reaction is in basic conditions, we now add OH^- ions to neutralize the H^+ ions. The same amount of OH^- is added to the other side.

In this case, we add $2OH^-$ to each side.

$$2OH^- + H_2O + N_2O + 2ClO^- \rightarrow 2NO_2^- + 2Cl^- + \overbrace{\cancel{2H^+} + \cancel{2OH^-}}^{2H_2O}$$

Step 9: Cancel out H_2O.

$$2OH^- + \cancel{H_2O} + N_2O + 2ClO^- \rightarrow 2NO_2^- + 2Cl^- + \overset{H_2O}{\cancel{2H_2O}}$$

$$2OH^- + N_2O + 2ClO^- \rightarrow 2NO_2^- + 2Cl^- + H_2O$$

Do a final check by adding up the charges on both sides to see if they balance:

$$\begin{array}{ccccccccccc} 2OH^- & + & N_2O & + & 2ClO^- & \rightarrow & 2NO_2^- & + & 2Cl^- & + & H_2O \\ (2 \times 1-) & & & & (2x1-) & & (2 \times 1-) & & (2 \times 1-) & & \\ 2- & + & 0 & + & 2- & = & 2- & + & 2- & + & 0 \end{array}$$

Both sides add up to 4-. This reaction is balanced.

Step 10: Add any spectator ions you removed in step 1.

(We didn't have any in this case)

Balanced Redox Reaction:

$$2OH^- + N_2O + 2ClO^- \rightarrow 2NO_2^- + 2Cl^- + H_2O$$

c) $NO_3^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow NO_{(g)} + ClO^-{}_{(aq)}$ (acidic conditions)

Step 1: Write the net ionic equation for the reaction and remove any spectator ions.

(In this case, this reaction is already written in net ionic form).

$$NO_3^- + Cl^- \rightarrow NO + ClO^-$$

Step 2: Write the reduction half-reaction and the oxidation half-reaction separately.

(HINT: If you are unsure which reaction is which you can assign oxidation numbers.)

Checking oxidation numbers	+5 2-		-1		+2 -2		+2 -2
	NO_3^-	+	Cl^-	→	NO	+	ClO^-

Oxidation: $Cl^- \rightarrow ClO^-$

Reduction: $NO_3^- \rightarrow NO$

Step 3: Balance all atoms **except any oxygen or hydrogen** (in this case they balance).

Oxidation: $Cl^- \rightarrow ClO^-$

Reduction: $NO_3^- \rightarrow NO$

Step 4: Balance any **oxygen first** by adding H_2O and then **hydrogen** by adding H^+

Oxidation: $H_2O + Cl^- \rightarrow ClO^- + 2H^+$

Reduction: $4H^+ + NO_3^- \rightarrow NO + 2H_2O$

Step 5: Total up the charge on each side of each half-reaction, and **add electrons (e^-)** to balance the charges.

$$\begin{array}{lcccccccccc}
\textit{Individual Charges:} & & 0 & & 1- & & 1- & & 2+ & & 2- \\
\text{Oxidation:} & & H_2O & + & Cl^- & \rightarrow & ClO^- & + & 2H^+ & + & 2e^- \\
\textit{Total charge:} & & & 1- & & = & & & 1- & &
\end{array}$$

$$\begin{array}{lccccccccc}
\textit{Individual Charges:} & 3- & & 4+ & & 1- & & 0 & & 0 \\
\text{Reduction:} & 3e^- & + & 4H^+ & + & NO_3^- & \rightarrow & NO & + & 2H_2O \\
\textit{Total charge:} & & & 0 & & & = & & 0 &
\end{array}$$

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 2 electrons and the reduction reaction has 3. We need to multiply the oxidation reaction by 3 and the reduction reaction by 2 in order to balance the electrons.

$$\begin{array}{ll}
\text{Oxidation:} & 3 \times [\ H_2O + Cl^- \rightarrow ClO^- + 2H^+ + 2e^-\ ] \\
 & 3H_2O + 3Cl^- \rightarrow 3ClO^- + 6H^+ + 6e^-
\end{array}$$

$$\begin{array}{ll}
\text{Reduction:} & 2 \times [\ 3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O\ ] \\
 & 6e^- + 8H^+ + 2NO_3^- \rightarrow 2NO + 4H_2O
\end{array}$$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons, H_2O, and H+ that appear on both sides.

$$\begin{array}{ll}
\text{Oxidation:} & \cancel{3H_2O} + 3Cl^- \rightarrow 3ClO^- + \cancel{6H^+} + \cancel{6e^-} \\
\text{Reduction:} & \cancel{6e^-} + \overset{2H^+}{\cancel{8H^+}} + 2NO_3^- \rightarrow 2NO + \overset{H_2O}{\cancel{4H_2O}} \\
 & 2H^+ + 2NO_3^- + 3Cl^- \rightarrow 2NO + 3ClO^- + H_2O
\end{array}$$

Do a final check by adding up the charges on both sides to see if they balance:

$$\begin{array}{ccccccccccc}
2H^+ & + & 2NO_3^- & + & 3Cl^- & \rightarrow & 2NO & + & 3ClO^- & + & H_2O \\
(2 \times 1+) & & (2 \times 1-) & & (3 \times 1-) & & & & (3 \times 1-) & & \\
2+ & + & 2- & + & 3- & = & 0 & + & 3- & + & 0
\end{array}$$

Both sides add up to 3-. This reaction is balanced.

Step 8: Add any spectator ions you removed in step 1.

(We didn’t have any in this case)

Balanced Redox Reaction:

$$2H^+ + 2NO_3^- + 3Cl^- \rightarrow 2NO + 3ClO^- + H_2O$$

d) $Cr^{3+}{}_{(aq)} + ClO^-{}_{(aq)} \rightarrow CrO_4^{2-}{}_{(aq)} + Cl^-{}_{(aq)}$ (basic conditions)

Step 1: Write the net ionic equation for the reaction and remove any spectator ions.

(In this case, this reaction is already written in net ionic form).

$$Cr^{3+} + ClO^- \rightarrow CrO_4^{2-} + Cl^-$$

Step 2: Write the reduction half-reaction and the oxidation half-reaction separately.

(HINT: If you are unsure which reaction is which you can assign oxidation numbers.)

Checking oxidation numbers

$$\begin{array}{ccccccc} +3 & & +1\ \ -2 & & +6\ \ -2 & & -1 \\ Cr^{3+} & + & ClO^- & \rightarrow & CrO_4^{2-} & + & Cl^- \end{array}$$

Oxidation: $Cr^{3+} \rightarrow CrO_4^{2-}$

Reduction: $ClO^- \rightarrow Cl^-$

Step 3: Balance all atoms **except any oxygen or hydrogen**.

In this case, the chromium and chlorine atoms are balanced:

Oxidation: $Cr^{3+} \rightarrow CrO_4^{2-}$

Reduction: $ClO^- \rightarrow Cl^-$

Step 4: Balance any **oxygen first** by adding H_2O and then **hydrogen** by adding H^+

Oxidation: $4H_2O + Cr^{3+} \rightarrow CrO_4^{2-} + 8H^+$

Reduction: $2H^+ + ClO^- \rightarrow Cl^- + H_2O$

Step 5: Total up the charge on each side of each half-reaction, and **add electrons (e^-)** to balance the charges.

$$\begin{array}{lccccccccc}
\textit{Individual Charges:} & & & & & & & (8 \times 1+) & & \\
 & 0 & & 3+ & & 2- & & 8+ & & 3- \\
\text{Oxidation:} & 4H_2O & + & Cr^{3+} & \rightarrow & CrO_4^{2-} & + & 8H^+ & + & 3\,e^- \\
\cline{2-4} \cline{6-10}
\textit{Total charge:} & & 3+ & & = & & & 3+ & &
\end{array}$$

$$\begin{array}{lccccccccc}
\textit{Individual Charges:} & & & (2 \times 1+) & & & & & & \\
 & 2- & & 2+ & & 1- & & 1- & & 0 \\
\text{Reduction:} & 2\,e^- & + & 2H^+ & + & ClO^- & \rightarrow & Cl^- & + & H_2O \\
\cline{2-6} \cline{8-10}
\textit{Total charge:} & & & 1- & & & = & & 1- &
\end{array}$$

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction has 3 electrons and the reduction reaction has 2. We need to multiply the oxidation reaction by 2 and the reduction reaction by 3 in order to balance the electrons.

$$\begin{array}{lll}
\text{Oxidation:} & 2 \times [& 4H_2O + Cr^{3+} \rightarrow CrO_4^{2-} + 8H^+ + 3\,e^- \;] \\
 & & 8H_2O + 2Cr^{3+} \rightarrow 2CrO_4^{2-} + 16H^+ + 6\,e^- \\
 & & \\
\text{Reduction:} & 3 \times [& 2\,e^- + 2H^+ + ClO^- \rightarrow Cl^- + H_2O \;] \\
 & & 6\,e^- + 6H^+ + 3ClO^- \rightarrow 3Cl^- + 3H_2O
\end{array}$$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons, water, and hydrogen ions.

$$\begin{array}{ll}
\text{Oxidation:} & \overset{5H_2O}{\cancel{8H_2O}} + 2Cr^{3+} \rightarrow 2CrO_4^{2-} + \overset{10H^+}{\cancel{16H^+}} + \cancel{6e^-} \\
\text{Reduction:} & \cancel{6e^-} + \cancel{6H^+} + 3ClO^- \rightarrow 3Cl^- + \cancel{3H_2O} \\
\hline
 & 5H_2O + 2Cr^{3+} + 3ClO^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 10H^+
\end{array}$$

Step 8: Because this reaction is in basic conditions, we now add OH^- ions to neutralize the H^+ ions. The same amount of OH^- is added to the other side.

In this case, we add 10 OH^- to each side.

$$10OH^- + 5H_2O + Cr^{3+} + 3ClO^- \rightarrow CrO_4^{2-} + 3Cl^- + \overbrace{\cancel{10H^+} + \cancel{10OH^-}}^{10H_2O}$$

Step 9: Cancel out H_2O.

$$10OH^- + \cancel{5H_2O} + 2Cr^{3+} + 3ClO^- \rightarrow 2CrO_4^{2-} + 3Cl^- + \overset{5H_2O}{\cancel{10H_2O}}$$

$$10OH^- + 2Cr^{3+} + 3ClO^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 5H_2O$$

Do a final check by adding up the charges on both sides to see if they balance:

$$\begin{array}{ccccccccccc} 10OH^- & + & 2Cr^{3+} & + & 3ClO^- & \rightarrow & 2CrO_4^{2-} & + & 3Cl^- & + & 5H_2O \\ (10 \text{ x } 1-) & & (2\text{x}3+) & & (3\text{x}1-) & & (2\text{x}2-) & & (3\text{x } 1-) & & \\ 10- & + & 6+ & + & 3- & = & 4- & + & 3- & + & 0 \end{array}$$

Both sides add up to 7-. This reaction is balanced.

Step 10: Add any spectator ions you removed in step 1.

(We didn't have any in this case)

Balanced Redox Reaction:

$$10OH^- + 2Cr^{3+} + 3ClO^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 5H_2O$$

e) $I_{2(s)} + Na_2S_2O_{3(aq)} \rightarrow Na_2S_2O_{4(aq)} + NaI_{(aq)}$ (acidic conditions)

Step 1: Write the net ionic equation for the reaction and remove any spectator ions.

(In this case, we need to write out a net ionic form.)

$I_2 + 2Na^+ + S_2O_3^{2-} \rightarrow 2Na^+ + S_2O_4^{2-} + Na^+ + I^-$ (full ionic equation)

$I_2 + \cancel{2Na^+} + S_2O_3^{2-} \rightarrow \cancel{2Na^+} + S_2O_4^{2-} + Na^+ + I^-$ (cancel spectator ions)

$I_2 + S_2O_3^{2-} \rightarrow S_2O_4^{2-} + I^-$ (net ionic equation)

Step 2: Write the reduction half-reaction and the oxidation half-reaction separately.

(HINT: If you are unsure which reaction is which you can assign oxidation numbers.)

Checking oxidation numbers

$$\begin{array}{ccccccc} 0 & & +2 \quad -2 & & 3 \quad -2 & & -1 \\ I_2 & + & S_2O_3^{2-} & \rightarrow & S_2O_4^{2-} & + & I^- \end{array}$$

Oxidation: $S_2O_3^{2-} \rightarrow S_2O_4^{2-}$

Reduction: $I_2 \rightarrow I^-$

Step 3: Balance all atoms **except any oxygen or hydrogen**. (In this case, we balance the iodine.)

Oxidation: $S_2O_3^{2-} \rightarrow S_2O_4^{2-}$

Reduction: $I_2 \rightarrow 2I^-$

Step 4: Balance any **oxygen first** by adding H_2O and then **hydrogen** by adding H^+

Oxidation: $H_2O + S_2O_3^{2-} \rightarrow S_2O_4^{2-} + 2H^+$

Reduction: $I_2 \rightarrow 2I^-$

Step 5: Total up the charge on each side of each half-reaction, and **add electrons (e^-)** to balance the charges.

Individual Charges:	0		2-		2-		2+		2-
Oxidation:	H_2O	+	$S_2O_3^{2-}$	$\rightarrow$	$S_2O_4^{2-}$	+	$2H^+$	+	$2e^-$
Total charge:		2-		=		2-			

Individual Charges:	2-		0		2-
Reduction:	$2e^-$	+	I2	$\rightarrow$	$2I^-$
Total charge:		2-		=	2-

Step 6: Multiply each half reaction by an appropriate number to balance the number of electrons between them.

In this case, the oxidation reaction and reduction reaction have the same number of electrons. There is no need to multiply either reaction.

Oxidation: $H_2O + S_2O_3^{2-} \rightarrow S_2O_4^{2-} + 2H^+ + 2e^-$

Reduction: $2e^- + I_2 \rightarrow 2I^-$

Step 7: Add the oxidation and reduction reactions to obtain an overall reaction. Cancel electrons, H_2O, and H+ that appear on both sides.

$$\begin{array}{lccccccccccccc}
\text{Oxidation:} & & & H_2O & + & S_2O_3^{2-} & \rightarrow & S_2O_4^{2-} & + & 2H^+ & + & \cancel{2e^-} \\
\text{Reduction:} & & & \cancel{2e^-} & + & I_2 & \rightarrow & 2I^- & & & & \\
\hline
 & H_2O & + & S_2O_3^{2-} & + & I_2 & \rightarrow & S_2O_4^{2-} & + & 2I^- & + & 2H^+
\end{array}$$

Do a final check by adding up the charges on both sides to see if they balance:

$$\begin{array}{ccccccccccc}
H_2O & + & S_2O_3^{2-} & + & I_2 & \rightarrow & S_2O_4^{2-} & + & 2I^- & + & 2H^+ \\
0 & + & 2- & + & 0 & = & 2- & + & 2- & + & 2+
\end{array}$$

Both sides add up to 2^-. This reaction is balanced.

Step 8: Add any spectator ions you removed in step 1.

(We now add back the sodium atoms. Refer back to the question.)

Balanced Redox Reaction:

$$H_2O + Na_2S_2O_3 + I_2 \rightarrow Na_2S_2O_4 + 2NaI + 2H^+$$

Notes

GRADE 12 CHEMISTRY (40S)

Module 2
Atomic Structure

Note to Students

Module 2 contains a number of images that are best viewed in colour. Colour versions of these images in PDF format are available in the learning management system (LMS).

MODULE 2:
ATOMIC STRUCTURE

Introduction to Module 2

The chemical properties of atoms are determined by their structure, more specifically by the arrangement of their electrons. In previous science courses, you learned about the structure of the atom, including the location of protons, neutrons, and electrons. Models such as the Bohr model do not sufficiently describe the observed properties of atoms. In this module, you will further examine the structure of atoms and relate their structure to their chemical properties.

Note

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.

Assignments in Module 2

Note

Please remember to work on Assignment 1.9: Researching a Redox Application as you work through this course. Please submit it at the end of Module 6.

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	Applications of Line Spectra
3	Assignment 2.2	Electron Configuration
4	Assignment 2.3	Valence Electrons
5	Assignment 2.4	Periodic Trends

LESSON 1: THE ELECTROMAGNETIC SPECTRUM (2 HOURS)

Lesson Focus

SLO C12-2-01: Describe qualitatively the electromagnetic spectrum in terms of frequency, wavelength, and energy.

SLO C12-2-02: Recognize, through direct observation, that elements have unique line spectra.
Include: flame tests or gas discharge tubes and spectroscopes or diffraction gratings

SLO C12-2-03: Describe applications and/or natural occurrences of line spectra.
Examples: astronomy, aurora borealis, fireworks, neon lights...

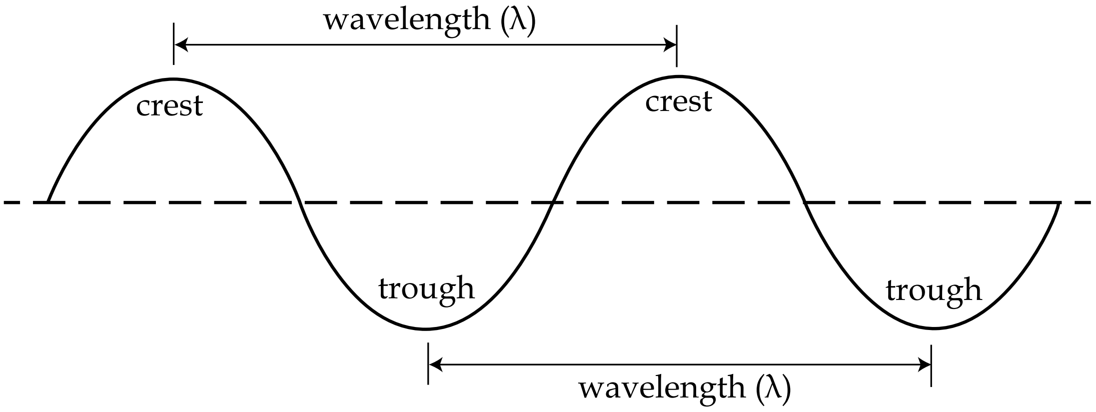
Lesson Introduction

To fully understand the behaviour of an atom, you must first understand the nature of light. This is because the light that you observe is actually generated by the movement of electrons. In the early 1900s, scientists observed that some substances emitted light when heated. This emitted light revealed information about the nature of the substance. In this lesson, you will learn more about light and how it relates to atomic structure.

Characteristics of Waves

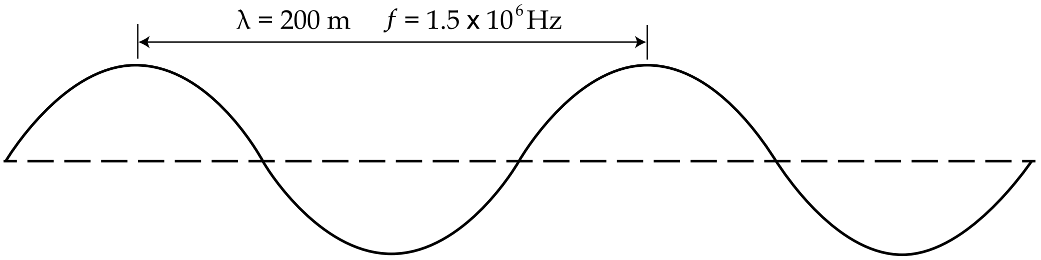
Since light travels in waves, it is useful to briefly introduce the characteristics of waves. The **wavelength** (symbol: λ, Greek letter “lambda”) is the distance from a point on one crest to the same point on the next crest. This distance is usually measured in metres (m). Another characteristic of waves is the frequency. **Frequency** (symbol: f) is the number of wavelengths, or wave cycles, that pass a point per unit of time. Frequency is measured in cycles per second, or the SI unit hertz (Hz). The frequency can also be represented by the unit s^{-1}, or the reciprocal of seconds.

Whether the waves are radio waves, red light waves, or X-rays, they all travel in a sinusoidal wave pattern, as the following diagram shows:

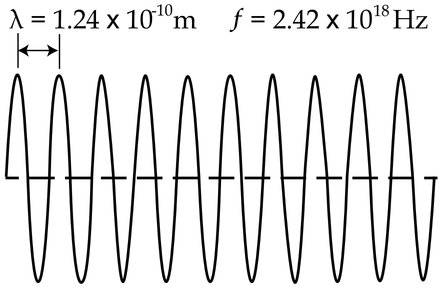


The following are two such examples.

Radio Wave



X-ray



Learning Activity 2.1: Wavelength and Frequency

What relationship do you notice between wavelength and frequency? Refer to the previous examples to help you answer this question.

Check the answer key.

Frequency, Wavelength, and Energy

In 1900, German physicist Max Planck (1858–1947), heated objects until they glowed, and then studied the light emitted by the glowing objects. He discovered that the frequency of this light was directly related to the amount of energy released.

Since the wavelength is inversely related to the frequency (in other words, as the wavelength increases, the frequency decreases and vice versa), the wavelength will also be inversely related to the energy of the light. That is, the greater the wavelength, the lower the energy. If you recall from the previous page, radio waves have a wavelength of about 2 m (frequency about 1.5 x 108 Hz) and X-rays have a wavelength of about 1.25 x 10^{-10} m (frequency about 2.4 x 10^{18} Hz). X-rays have a greater amount of energy than radio waves. Consequently, radio waves are all around us (you can get a radio signal almost anywhere) and are not harmful to us, whereas exposure to X-rays is limited because their high energy is damaging.

The Electromagnetic Spectrum

If you shine sunlight through a prism, you observe a rainbow of colours known as a **spectrum**.

Remember that these and other images are available in colour in the learning management system (LMS).

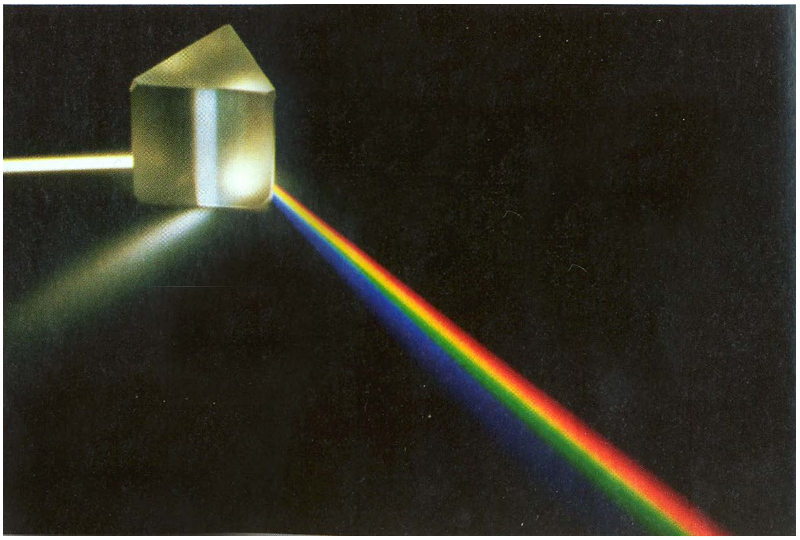
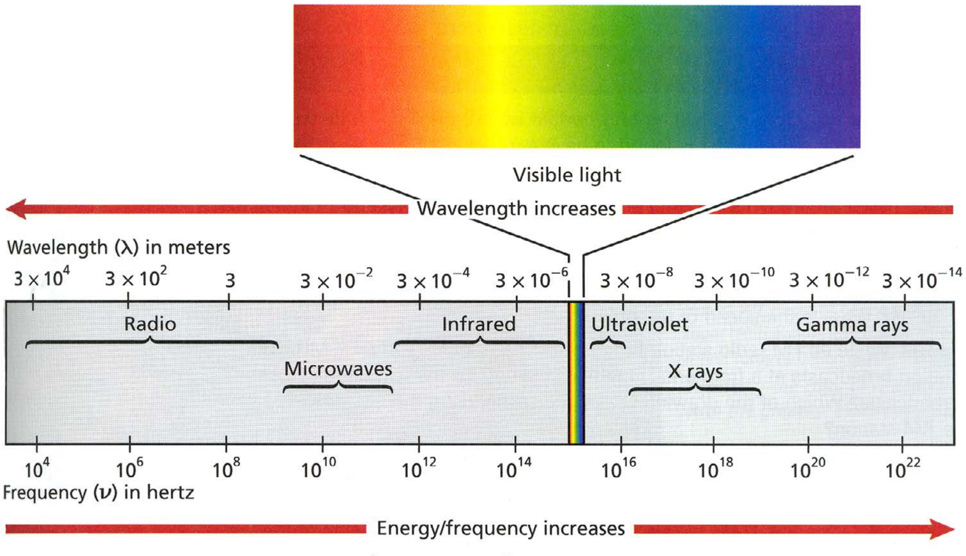


Photo Credit: Courtesy of David Parker/Science Photo Library/Photo Researchers
Buthelezi, Thandi, *et al. Chemistry: Matter and Change: Student Text.* New York, USA: The McGraw-Hill Companies, Inc., Glencoe Science, 2008. 138.

Each colour in the rainbow represents light of a different frequency, or wavelength. This is known as a continuous spectrum because there are no breaks between the individual wavelengths. Our eyes perceive this spectrum as white. The visible spectrum is actually a very small portion of the entire electromagnetic spectrum and is found at a wavelength of approximately 10^{-6} m.



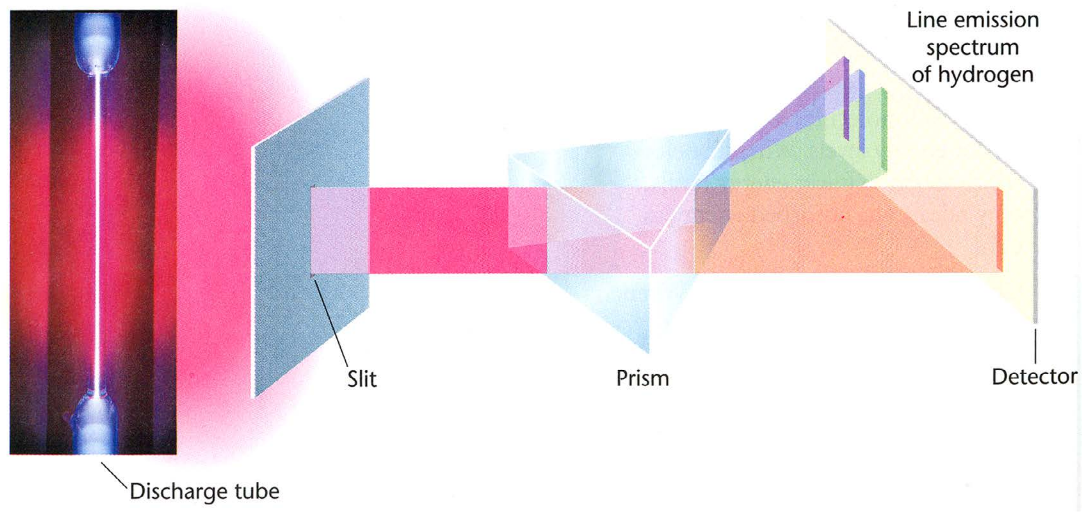
Electromagnetic Spectrum

Buthelezi, Thandi, *et al. Chemistry: Matter and Change: Student Text.* New York, USA: The McGraw-Hill Companies, Inc., Glencoe Science, 2008. 139.

James Clerk Maxwell determined that light is formed by electrical and magnetic waves moving perpendicularly to one another. For this reason, light is considered a part of the electromagnetic spectrum.

Line Spectra

When an electric current passes through hydrogen gas in a tube, the gas glows. If the light produced by the glowing gas is focused through a slit and then passed through a prism, a spectrum with distinct lines is produced.



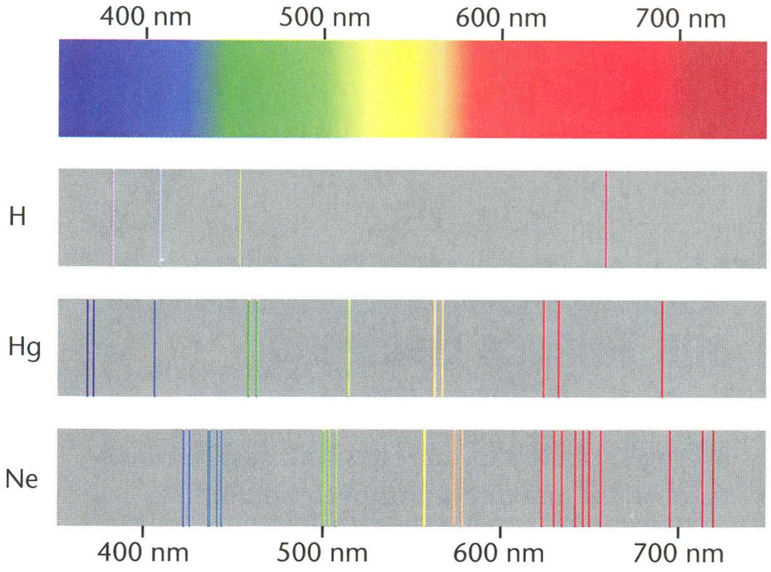
Emission Spectrum of Hydrogen
When hydrogen atoms are energized by electricity, they emit light. ▼

A prism will separate the light emitted by hydrogen into four visible lines of specific frequencies (colors). Hydrogen also emits light in the non-visible ultraviolet and infrared parts of the spectrum. ▼

Photo Credit: Courtesy of Rich Treptow/Photo Researchers
Phillips, John S., *et al. Chemistry: Concepts and Applications.* New York, USA: The McGraw-Hill Companies Inc., Glencoe/McGraw-Hill, 2000. 74.

This type of spectrum is known as an emission spectrum because it is, in fact, the separate wavelengths of light emitted by the gas. This emission spectrum is also known as a line spectrum, because the light separates into discrete wavelengths of light that appear as lines of colour on a screen or photographic plate. Unlike a continuous spectrum, the colours in a line spectrum do not blend into each other. The visible region of the emission spectrum of hydrogen is shown at www.mhhe.com/physsci/chemistry/essentialchemistry/flash/linesp16.swf. Each coloured line corresponds to an exact wavelength, or frequency, of light.

Spectroscopy, **spectrophotometry**, and **spectrometry** are several names given to the techniques used to determine a substance's emission spectrum. The following are the emission or line spectra of H (hydrogen), Hg (mercury), and Ne (neon):



Phillips, John S., *et al. Chemistry: Concepts and Applications.* New York, USA: The McGraw-Hill Companies Inc., Glencoe/McGraw-Hill, 2000. 235.

If you examine these spectra, you will notice that they are all different. The line spectrum of each element is like the fingerprint of that element.

Flame Tests

When some elements are burned, they give off a distinctive colour of light that is characteristic of that element. This is caused by the emission of light of specific wavelengths in the line spectrum. Preliminary evidence of the presence of a metal is often the colour that results when it is placed in a flame. Salts containing metals will give off characteristic colours when placed into a flame. This is known as a flame test.

Some commercially made fireplace logs advertise that they will produce multicoloured flames when burned. This is because the logs are made of pressed sawdust or smaller chips of wood soaked in a solution containing several metal salts. You can also see colourful flames if you burn the paper from newspapers or magazines with colour pictures, since the coloured inks contain various metal salts.

In the video clip at www.dlt.ncssm.edu/core/c8.htm, you can observe the distinct and characteristic colours that different metal ions produce.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Applications of Line Spectra

Where can you observe line spectra in the world around you? The aurora borealis (northern lights) are a result of atmospheric gases in an excited state (due to electrons being raised to higher energy levels). This excitation leads to the release of light as electrons return to a lower energy level.

Another familiar application of line spectra is in the manufacturing of fireworks. You saw the distinctive colours produced by metal ions in the animation of flame tests. These metal salts are used in fireworks to produce the distinctive colours of light that you see burst in the sky.

Learning Activity 2.2: Line Spectra

1. Put the following in order of increasing energy: green light, X-rays, radio, red light, ultraviolet, microwaves, blue light, and gamma rays.
2. Describe the relationship between frequency, wavelength, and energy.
3. Why is light called electromagnetic radiation?
4. Compare and contrast line spectra and continuous spectra.
5. Using the flame test video clip, complete the following table:

Metal Ion	Flame Colour
Potassium	
Barium	
Calcium	
Strontium	
Lithium	
Sodium	
Copper	

Check the answer key.

Lesson Summary

In this lesson, you learned that light travels in waves. Wavelength and frequency are inversely related. In other words, as wavelength increases, the frequency decreases. You observed examples of light spectrum, and you saw that each visible colour represents light of a different frequency, or wavelength. This is known as a continuous spectrum because there are no breaks between the individual wavelengths: the colours blend together. The visible spectrum is actually a very small portion of the entire electromagnetic spectrum.

Notes

Assignment 2.1: Applications of Line Spectra (10 marks)

Choose one application or natural occurrence of line spectra (e.g., astronomy, aurora borealis, fireworks, neon lights).

Use the Internet, magazines, books, or other resources to help you find information. Answer the following questions:

1. Describe the application(s) or natural occurrence of line spectra that you chose. *(2 marks)*

2. What elements are used/profiled in your example? List the colours they produce. *(4 marks)*

continued

Assignment 2.1: Applications of Line Spectra (continued)

3. Describe the line spectra in terms of the wavelength and frequency for the colour(s) you would see. *(2 marks)*

4. Provide a list of information sources that you used in this assignment. *(2 marks)*

LESSON 2: THE QUANTUM MECHANICAL MODEL (3 HOURS)

Lesson Focus

SLO C12-2-04: Outline the historical development of the quantum mechanical model of the atom.

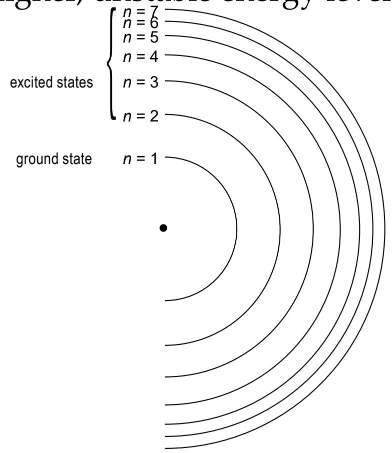
Lesson Introduction

In the previous lesson, you learned about light waves and the emission spectrum. You are now aware that continuous spectrum contains all wavelengths of light and that line spectrum has distinct wavelengths of light. In this lesson, you will continue your study of atomic structure, outlining the historical development of the quantum mechanical model of the atom.

Neils Bohr Explains Line Spectra

In 1922, **Neils Bohr** (1885–1962) developed a model that explained hydrogen's line spectra. His model dealt specifically with the movement of electrons around the atom. Bohr proposed that the electrons in a hydrogen atom are arranged in stable orbits around the nucleus depending on their energy. He said the orbits were like a ladder with unequally spaced rungs, or a stairway with unequal stairs.

When radiation is absorbed by an atom, an electron may gain energy and jump from its normal orbit, also called the **ground state** (its resting state), to a higher, unstable energy level called the **excited state**.



Adapted from Spence, D. "Principal Energy Levels and Sub-orbitals." *Mr. Spence's Science Website*. (21 Feb. 2013). http://sun.menloschool.org/~dspence/chemistry/atomic/config2.html.

The model then states that this electron eventually loses energy and reverts to a lower energy level by emitting energy in the form of light. The individual lines of the hydrogen line spectra are due to the hydrogen electron becoming excited, moving to a higher orbit, and then falling to a lower orbit and releasing energy in the form of light. The distinct lines are due to the distinct orbits or energy levels the electron can occupy.

An animation showing the electron in a hydrogen atom absorbing energy and moving to its excited state is available online at www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/ElectronOrbits.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem1.htm#atomic
2. scrolling down the page until you find the animation titled "**ElectronOrbits.html**"

When the hydrogen moves back to its former ground state, a photon of light is emitted.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

The Quantum Mechanical Model of the Atom

In Grade 9 and Grade 10 Science, you relied on the Bohr model to base assumptions you made about the behaviour of an atom. **Neils Bohr** made a significant contribution to our understanding of atoms, but his theory did not provide a complete description of electronic behaviour in the atom. Despite his work, one question regarding the position of the electron remained unanswered: If an electron moves in a wave-like motion, how can we predict its precise location within the atom?

This led to the development of the Heisenberg uncertainty principle by **Werner Heisenberg** (1901–1976). According to Heisenberg, it is impossible to know simultaneously both the momentum (or movement) and the position of an electron with certainty. Essentially, it is impossible to simultaneously know both the exact location of an electron as well as where it is going, because in order to "see" the electron we must shine a light on it (we can't see in the dark). Shining a light on the electron changes its motion because the photons (light particles) hit the electron and so we can't definitively predict where it was or where it was going.

In 1926, **Erwin Schrödinger** (1887–1961) used complex math to develop an equation that describes the energies and behaviour of tiny particles such as electrons. This equation represents the statistical probability of finding an electron in a particular space in the atom. It began a new era in physics and chemistry as it launched the new field of **quantum mechanics**.

The foundation of quantum mechanics is the idea that energy comes in bundles which can’t be further divided. A quantum of energy refers to the small packets of energy that can be absorbed or radiated but can’t be divided.

Principal Energy Levels

It is at this point in history that further elaboration could be made to the Bohr model.

According to Schrödinger’s equation, Bohr’s orbits or energy levels become **principal quantum numbers** (*n*), also called principal energy levels. Within each energy level, there is a set of **orbitals** or **sublevels**.

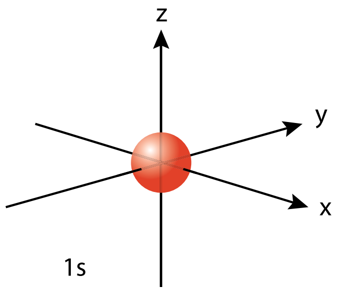
In Bohr’s idea of the atom, the electrons move in orbits at specified distances based on their energy. Electrons with higher energy are in orbits further away from the nucleus. Schrödinger’s work built upon Bohr’s model by determining that the electrons move more or less randomly around the nucleus, but that their energy level determines the region around the nucleus in which you will likely find them. Bohr’s orbits are more like boundaries—electrons are found, most of the time, within the boundary dictated by their energy.

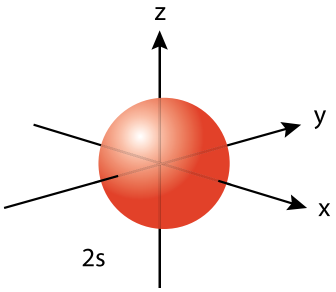
The lowest energy level is $n = 1$ and the highest is $n = 7$. As the value of n increases, the energy levels become larger and the electrons spend more time further from the nucleus.

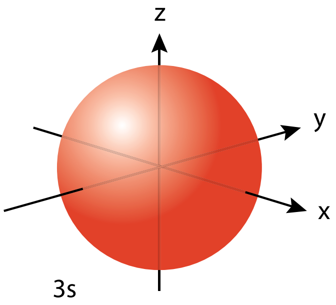
Electron Orbitals

You have likely learned that electrons fill the space around the nucleus of an atom. In an attempt to visualize this, different representations of electron orbitals have been proposed. **Electron orbitals** are three-dimensional regions where there is a 90% probability of finding electrons. (This means that there is a 10% probability that the electron could be almost anywhere!) The orbitals are described in terms of three variables: size, shape, and orientation in space. There are five types of orbitals with different regions of probability, given the letter names *s*, *p*, *d*, *f*, and *g*. These regions of probability result in fuzzy electron clouds that have different shapes.

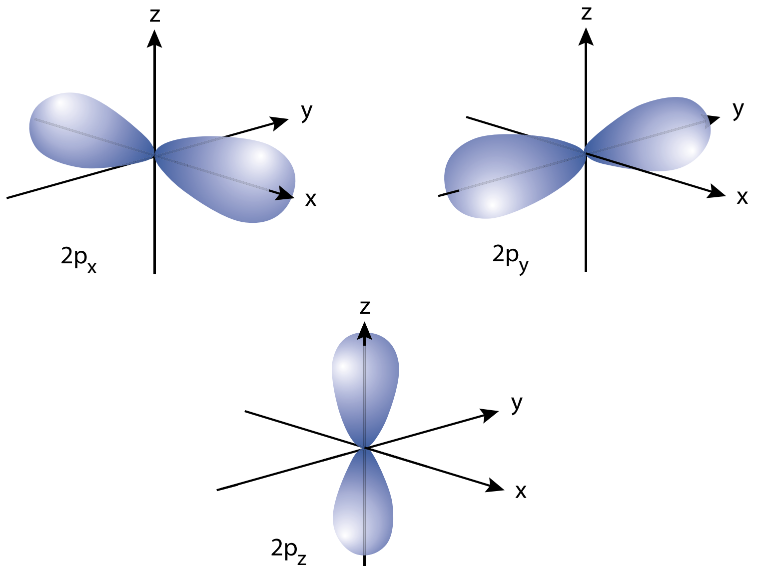
- The **s-orbitals** are spherical in shape. Each energy level has one s-orbital. The s-orbital in the first energy level is given the designation 1s, the s-orbital in the second energy level has the designation 2s, and so on. Following are representations of the s-orbitals for the first three energy levels. Notice how the size of the orbital increases as the principal quantum number, or energy level, increases. The position of the nucleus is assumed to be the point where the three axes intersect.



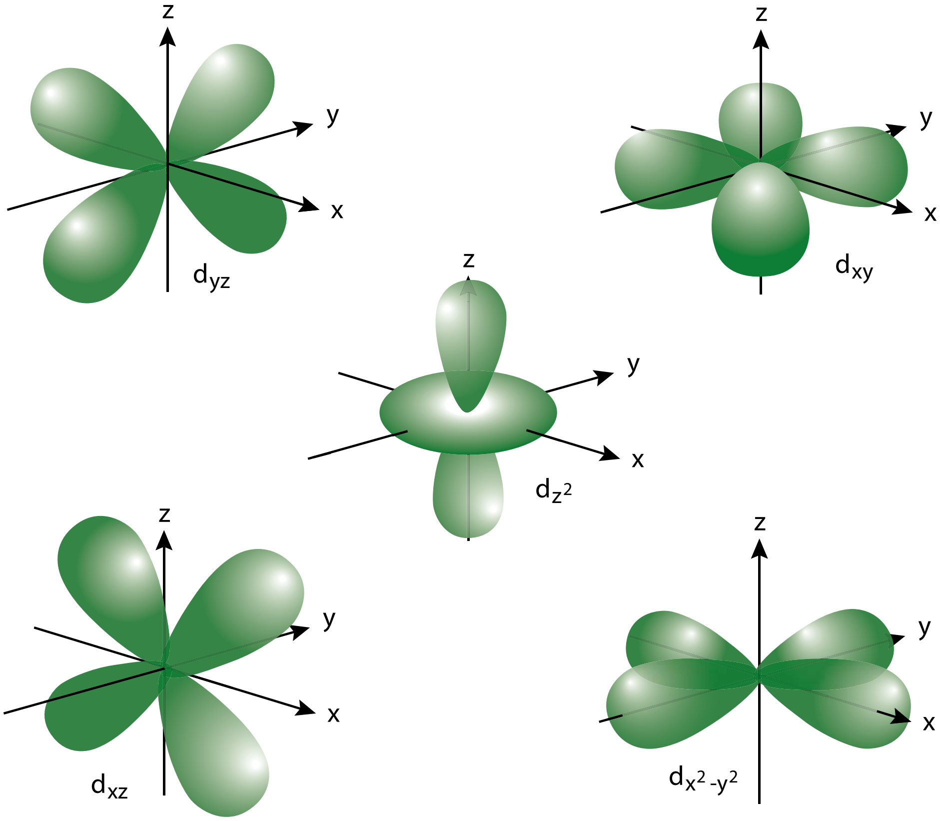




- The **p-orbitals** have two regions, or lobes, of high probability. Since these are regions of probability, electrons spend equal amounts of time in both lobes of the p-orbital. The p-orbitals are only present in the second energy level and higher. There are three kinds of p-orbitals, depending on their position in space (along the x-axis, along the y-axis, or along the z-axis).



- The **d-orbitals** are only present in the third energy level and up. They have varied shapes and occupy more complex regions in space than the p-orbitals. There are five types of d-orbitals.



- The **f-orbitals** are only present in the fourth energy level and up. There are seven different types of f-orbitals and their shapes are considerably more complex than the d-orbitals (some containing as many as eight lobes).
- The **g-orbitals** occur only in energy levels 5 and up. For this course, you only need to be familiar with the s, p, and d orbitals up to energy level 4. The very complex orbitals found in g will not be covered in this course.

If you would like to learn more about electron orbitals, visit the Orbitron website, a gallery of atomic orbitals and molecular orbitals, at http://winter.group.shef.ac.uk/orbitron/.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Learning Activity 2.3: The Quantum Mechanical Model

1. Identify the contribution made to atomic theory by each of the following scientists:
   a. Bohr
   b. Schrödinger
2. What is the Heisenberg uncertainty principle?
3. Complete the following table describing the s-, p-, and d-orbitals.

Orbital	**s-orbital**	**p-orbital**	**d-orbital**
Shape			
Number of orbitals in each principal energy level			
Found in which principal energy levels			

4. What are orbitals?
5. Compare ground state and excited state.
6. How many orbitals (or sublevels) are found in the fourth principal energy level?
7. Complete the following sentences with the correct term.
   a. According to the Heisenberg uncertainty principle, it is impossible to know the ____________________________ and the position of a particle at the same time.
   b. Energy is ____________________________ when an electron moves from ground state to excited state.

continued

Learning Activity 2.3: The Quantum Mechanical Model (continued)

c. Electron ____________________________ are areas of high probability of finding electrons.

d. Orbitals are described in terms of size, shape, and ____________________________ in space.

e. Principal quantum numbers, also called principal energy levels, are represented by the symbol ____________________________.

f. The size of the s-orbital ____________________________ as the principal quantum number, or energy level, increases.

Check the answer key.

Lesson Summary

In this lesson, you learned that there are representations of electron orbitals other than those you may have previously known. These electron orbitals are regions in the atom of high probability of finding electrons. Electron orbitals are described in terms of three variables: size, shape, and orientation in space. In the next lesson, you will learn about electron configurations and relate them to electron orbitals.

LESSON 3: ELECTRON CONFIGURATION (3 HOURS)

Lesson Focus

SLO C12-2-05: Write electron configurations for the elements of the periodic table.
Include: selected elements up to atomic number 36 (krypton)

Lesson Introduction

In the previous lesson, you learned about the quantum mechanical model of the atom. Using this model, scientists can predict the location of electrons according to principal energy levels and sublevels. The characteristics of elements are due to this arrangement of their electrons. In this lesson, you will examine the arrangement of the electrons around the nucleus and predict some of the properties of that element.

Electron Orbitals: A Review

In the previous lesson, you learned that orbitals are regions of space around the nucleus of the atom where there is a high probability of finding an electron.

You also saw that each orbital is given the letter designation of *s, p, d,* or *f*. Each of these orbitals has a different, characteristic shape. There are also a different number of each type of orbital possible in each energy level, *n*. For example, there is one s-orbital in every energy level, but there are three p-orbitals in each sublevel. The total number of orbitals in each energy level can be calculated by squaring the principal number (n^2). The table below summarizes the types and number of orbitals available in each energy level.

Principal Energy Level or Principal Quantum Number (n)	Number of Orbitals (n^2)	Orbital Types
1	1	1s-orbital
2	4	1s-orbital + 3p-orbitals
3	9	1s-orbital + 3p-orbitals + 5d-orbitals
4	16	1s-orbital + 3p-orbitals + 5d-orbitals + 7f-orbitals

The Pauli Exclusion Principle

Did you know that electrons spin on their own axes much like Earth? Electrons are constantly spinning, and this spinning produces a magnetic field. According to **Wolfgang Pauli**, two identical electrons cannot occupy the same orbital. This makes sense because electrons having the same negative charge repel each other. Pauli said that two electrons can only occupy the same orbital if they have **opposite spins**. These opposite spins create a north and south pole, like mini-magnets. This magnetic field helps each electron overcome the other's repulsive forces, allowing two electrons to occupy the same orbital.

An animation showing the opposite spin of two electrons that are occupying the same orbital is available online at www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/ElectronSpin.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem1.htm#atomic
2. scrolling down the page until you find the animation titled "**ElectronSpin.html**" or "**ElectronSpin.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Knowing that the maximum number of electrons per orbital is two, then the number of electrons per energy level will be 2 × n^2 (or $2n^2$). The table you saw in the previous section can now be expanded to include the number of electrons in each energy level.

Principal Energy Level or Principal Quantum Number (n)	**Number of Orbitals (n^2)**	**Orbital Types**	**Number of Electrons ($2n^2$)**
1	1	1×s	2
2	4	1×s + 3×p	8
3	9	1×s + 3×p + 5×d	18
4	16	1×s + 3×p + 5×d + 7×f	32

Electron Configuration

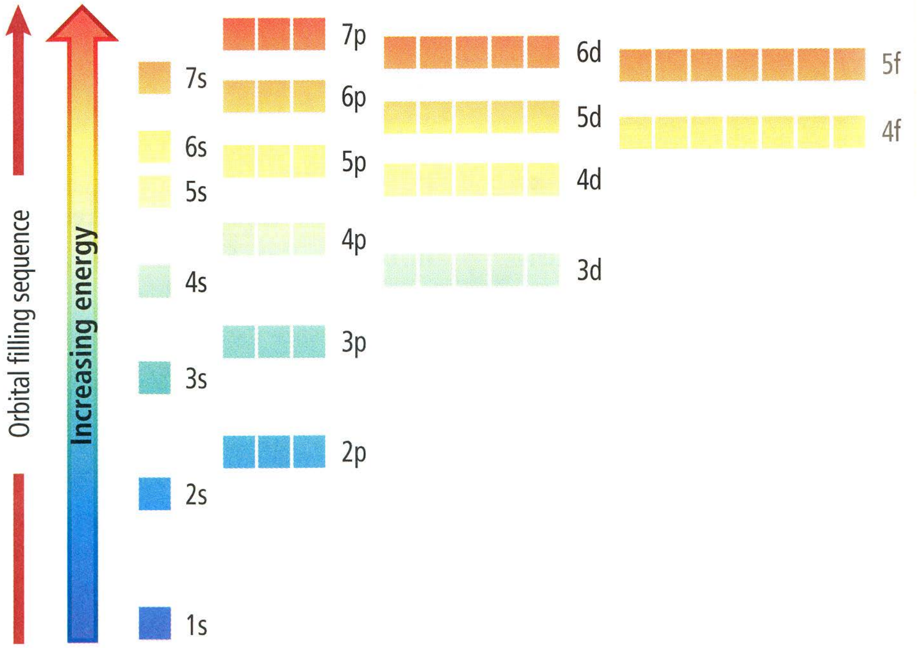
The specific arrangement of electrons in orbitals around a nucleus is called **electron configuration**. A configuration is an arrangement of parts; in this case, an arrangement of electrons within the orbitals.

You can use the following three rules to help you determine the electron configuration of a given atom.

1. The Aufbau Principle

The **Aufbau principle** (aufbau is German for "buildup") was proposed by Neils Bohr. According to this principle, electrons must be added to the lowest energy state before being added to a higher energy level. In each energy level, the s-orbital has the lowest energy, therefore, it is always the first orbital filled in that energy level. The next energy level is the p-orbital.

Note, however, **that the 4s-orbital happens to have a lower energy than the 3d-orbitals**. The reason for this overlap is beyond the scope of this course. Also in the Aufbau principle, electrons will fill the 4s-orbital before filling the 3d-orbitals. When the 3d-orbitals are full, electrons begin filling the 4p-orbitals. The same trend of s orbitals of the next level filling before the d-orbitals of the lower level occurs through to the eighth period.



Buthelezi, Thandi, *et al. Chemistry: Matter and Change: Student Text.* New York, USA: The McGraw-Hill Companies, Inc., Glencoe Science, 2008. 156.

2. The Pauli Exclusion Principle

This principle states that each orbital gets two electrons. According to the **Pauli exclusion principle**, each orbital can hold two electrons, providing they have opposite spins. These electrons are represented by arrows, and the arrowheads indicate the direction of the electron spin. There is only one electron in the hydrogen atom, which we arbitrarily choose to represent with an upward arrow. The direction you begin with is not important.

The helium atom has two electrons to fill the s-orbital in the first principal energy level. The electron must fill the orbital with an opposite spin, represented by an upward and downward arrow. This trend continues until we reach the carbon atom.

3. Hund's Rule

Friedrich Hund (1896–1997) proposed that in orbitals that have several suborbitals (p, d, f, etc.), electrons fill all the suborbitals with single electrons before filling them with second electrons. A common analogy is people boarding a bus: each seat can hold two people; however, people usually sit by themselves and only double up once all the seats have one person. Hund also stated that the first electrons in the suborbitals all have the same spin and the second electrons added to the suborbitals have the reverse spin. This is called **Hund's rule**.

Below is a table that includes both the orbital diagrams and the electron configurations for the first 10 elements. An **orbital diagram** represents the spin of an electron in terms of up (↑) and down (↓) arrows. These spins can be thought of "up" and "down," or "clockwise" and "counterclockwise."

Element	Filling of Orbitals					Electron Configuration
	1s	2s	$2p_x$	$2p_y$	$2p_z$	
H	↑					1s
He	↑↓					$1s^2$
Li	↑↓	↑				$1s^22s^1$
B	↑↓	↑↓	↑			$1s^22s^22p^1$
C	↑↓	↑↓	↑	↑		$1s^22s^22p^2$
N	↑↓	↑↓	↑	↑	↑	$1s^22s^22p^3$
O	↑↓	↑↓	↑↓	↑	↑	$1s^22s^22p^4$
F	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^22s^22p^6$

Note that when the p-orbitals fill from boron (B) through to neon (Ne), the $2p_X$, $2p_Y$, and $2p_Z$ orbitals receive one electron of the same spin each before the second, opposite-spin electrons are added.

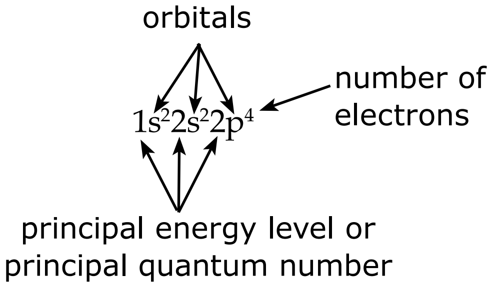
Writing Electron Configuration

Now that you have learned some important basics, it is time to try writing electron configurations for some atoms. The electron configuration of an atom assigns a number for each principal energy level, as well as a number to each electron.

Note

- The number in the coefficient position refers to the principal energy level of the electron.
- The lowercase letter refers to the name of the orbital.
- The superscript number refers to the number of electron(s) in that orbital.

For example, the electron configuration of oxygen is as follows:



The electron configuration above shows that there are two electrons in the s-orbital of the first energy level ($1s^2$), two electrons in the s-orbital of the second energy level ($2s^2$), and four electrons in the p-orbital of the second energy level ($2p^4$).

Note

The orbitals are always written as lowercase letters, and the number of electrons is always written in superscript.

Example 1

Write the complete electron configuration for magnesium.

- The atomic number = 12; therefore, the number of electrons = 12.
- Follow the periodic table (or use the Aufbau principle), adding electrons until you reach 12 electrons.

$1s^2 2s^2 2p^6 3s^2$

Example 2

Write the complete electron configuration for germanium.

- The atomic number is 32; accordingly, the number of electrons is 32.
- Follow the periodic table (or use the Aufbau principle), adding electrons until you reach 32 electrons. Remember the 4s-orbital fills first due to overlapping.

  $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$

Noble Gas Configuration

On occasion, electron configurations can get very long. In such a case, you may choose to write the **noble gas configuration** for that element. This is also called a **kernel** (the highest noble gas structure within the configuration). Below is the kernel configuration from Example 2. In this case, the highest noble gas structure within the configuration is argon. We then write the symbol for argon in square brackets followed by the remaining configuration.

$[Ar]4s^23d^{10}4p^2$

Example

What is the electron configuration for nickel in terms of the nearest noble gas?

$[Ar]\ 4s^2\ 3d^8$

How would the last valence orbital be filled?

3d

↑↓	↑↓	↑↓	↑	↑

This **kernel configuration** is helpful when writing the configurations of atoms containing many electrons.

Exceptions to the Rules

Not all elements have the electron configuration you would expect. There are certain cases where the orbitals do not fill in the order you have just learned. You may want to memorize the following examples in case they come up in an assignment or examination.

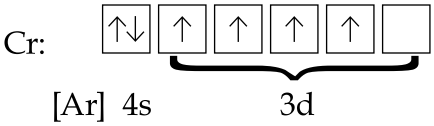
Chromium, copper, and their families undergo what is called electron promotion.

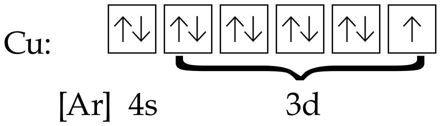
If you follow the Aufbau principle, the electron configurations for chromium and copper should be as follows:

Cr: $1s^22s^22p^63s^23p^64s^23d^4$ or $[Ar]4s^23d^4$

Cu: $1s^22s^22p^63s^223p^64s^23d^9$ or $[Ar]4s^23d^9$

Or, you can look at the orbital diagrams for the electrons outside the kernel.



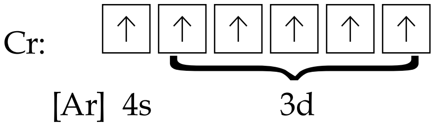


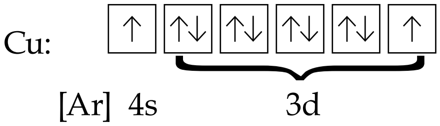
There is, however, increased stability of full or half-filled orbitals that causes the electron promotion. In the above examples, electrons from the s-orbital will actually move to the d-orbital to achieve this more stable, half-full state. As a result, the actual configurations for Cr and Cu are as follows:

Cr: $1s^22s^22p^63s^23p^6\mathbf{4s^13d^5}$ or $[Ar]\mathbf{4s^13d^5}$

Cu: $1s^22s^22p^63s^23p^6\mathbf{4s^13d^{10}}$ or $[Ar]\mathbf{4s^13d^{10}}$

The orbital diagrams illustrate why this configuration is an advantage.





In chromium, the promotion of a 4s-electron to the empty 3d-orbital creates half-filled orbitals. In copper, using a 4s-electron to fill the 3d-orbital is a more stable configuration. At higher energy levels, the exceptions to the Aufbau principle increase because the energy differences between sublevels become smaller. This means that there are many more exceptions when filling the 5d- and 4f-orbitals.

Learning Activity 2.4: The Aufbau Principle

1. How many electrons in an atom can be found in the following orbital, sublevel, or principal energy level?
   b. 3d ____________________
   c. n=2 ____________________
   d. 4p ____________________
   e. f ____________________
   f. s ____________________
   g. 2p ____________________
   h. 3d ____________________
   i. n=5 ____________________
2. Write complete electronic configurations for the following atoms and ions:
   a. Cu ____________________
   b. Zn ____________________
   c. P ____________________
   d. Ca ____________________
3. How many unpaired electrons are there in each of the following?
   a. Ca ____________________
   b. Ge ____________________
   c. Li ____________________
   d. O ____________________
   e. Mn ____________________
   f. P ____________________

Check the answer key.

Lesson Summary

In this lesson, you learned three rules to help you determine the electron configuration of a given atom. You used your knowledge of an element's position on the periodic table to help you write electron configurations. In the next lesson, you will relate the electron configuration of an element to its valence electron(s) and its position on the periodic table.

Notes

Assignment 2.2: Electron Configuration (12 marks)

1. Complete electronic configurations for the following atoms. *(2 marks each)*

Element	# of Electrons in Element	Electron Configuration
He	2	
Li	3	
Be	4	
O	8	
Cl	17	
K	19	

Notes

LESSON 4: VALENCE ELECTRON CONFIGURATION (1 HOUR)

Lesson Focus

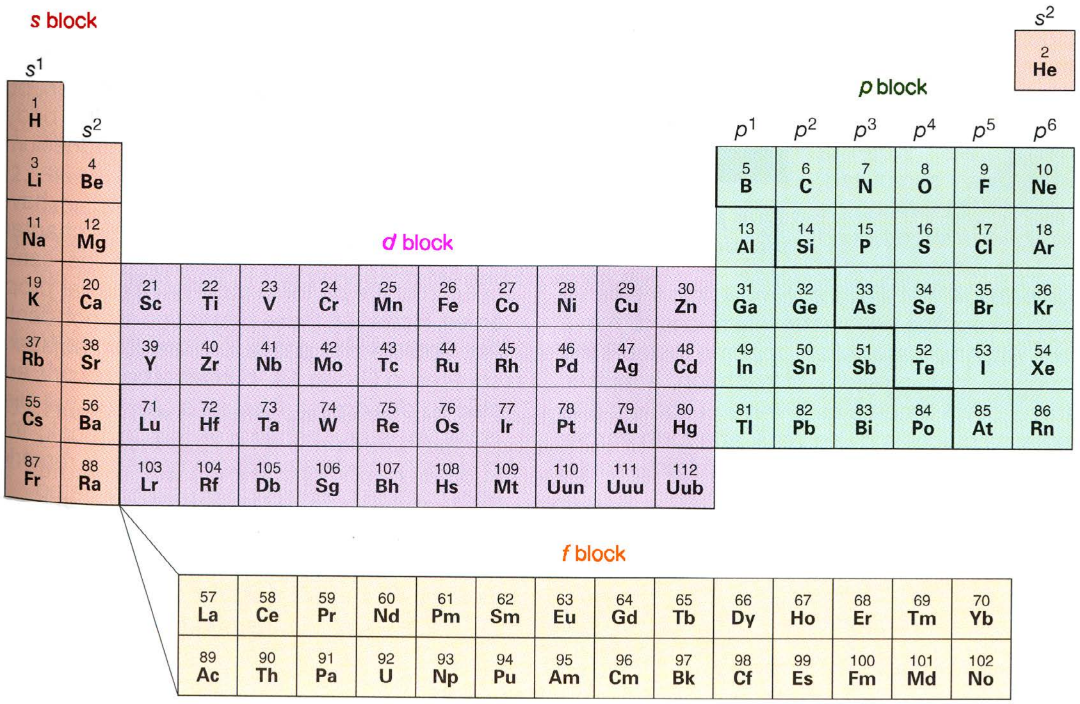
SLO C12-2-06: Relate the electron configuration of an element to its valence electron(s) and its position on the periodic table.

Lesson Introduction

In the previous lesson, you learned about electron configuration and how the configuration of an element relates to its position on the periodic table. In this lesson, you will further relate the valence electron configuration of an element to its position on the periodic table.

Orbitals relate to the periodic table in a very direct way. You already know that the period number is the same as the principal energy level. The following are more connections between the periodic table and electron configuration:

- The alkali metals and alkali earth metals are the two families that make up the s-block.
- The six families from group 13 (the Boron family) to group 18 (the noble gases) make up the p-block.
- The transition metals from group 3 to group 12 make up the d-block.
- The actinides and the lanthinides make up the f-block.



LeMay, H. Eugene, Jr., *et al. Chemistry: Connections to Our Changing World.* 2nd ed. Upper Saddle River, NJ: Prentice-Hall Inc., 2000. 171.

Using the labelled periodic table above, you can relate the electron configuration of an atom to its position on the periodic table. For example, the electron configuration for the element nickel (28) would look like this:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$

Following from left to right along the periods, you might write the electron configuration for niobium (41) like this:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^3$

However, the correct configuration is actually:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ \mathbf{5s^1\ 4d^4}$

Don’t forget that electrons will fill the 5s-orbital before filling the 4d-orbitals!

Valence Electrons

The electronic configuration of an atom is given by listing its subshells with the number of electrons in each subshell. Remember that electrons are added to the subshell of lowest energy until it reaches its capacity, at which time the subshell of the next energy level begins to be filled. The electrons on the highest numbered subshells are the **valence electrons**, which comprise the valence shell of the atom. It is the valence electrons that are lost or gained in chemical reactions, and these electrons will always be from the s- and p-orbitals. See the table below for some examples that relate expanded electron configuration and valence electron configurations.

Element Name	Electron Configuration	Valence Shell
Hydrogen	**$1s^1$**	$1s^1$
Neon	$1s^2$**$2s^22p^6$**	$2s^22p^6$
Phosphorus	$1s^22s^22p^6$**$3s^23p^3$**	$3s^23p^3$
Calcium	$1s^22s^22p^63s^23p^6$**$4s^2$**	$4s^2$

Often, you will see only the valence electron configuration of an atom, like $3s^1$ for sodium or $4s^2$ for calcium. These electrons, called **valence electrons**, are the most distant from the positive nucleus and are transferred between atoms in chemical reactions. In the previous table, the valence electrons are in bold.

The electron configuration of an element is related both to its valence electrons and to its position on the periodic table. Remember the following:

- The alkali metals have 1 valence electron (s^1).
- The alkali earth metals have 2 valence electrons (s^2).
- The boron family has 3 valence electrons (s^2p^1).
- The carbon family has 4 valence electrons (s^2p^2).
- The nitrogen family has 5 valence electrons (s^2p^3).
- The chalcogen family has 6 valence electrons (s^2p^4).
- The halogen family has 7 valence electrons (s^2p^5).
- The noble gases have 8 valence electrons (s^2p^6).

Electron Configurations of Ions

You have already learned that an ion is the result of either losing the valence electrons (leaving a positive ion) or gaining enough electrons to fill that valence subshell. As such, the ion configuration reflects the new number of electrons within that ion. The table below shows the relationship between loss and gain of electrons and ion configuration.

Chemical Element	Valence Shell	Electron Transfer	Resulting Ion	Ion Configuration
Cl	$3s^2 3p^5$	gain 1	Cl^-	$1s^2 2s^2 2p^6 3s^2 3p^6$
Ar	$3s^2 3p^6$	none	Ar^0	$1s^2 2s^2 2p^6 3s^2 3p^6$
K	$4s^1$	lose 1	K^+	$1s^2 2s^2 2p^6 3s^2 3p^6$
Ca	$4s^2$	lose 2	Ca^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6$

Did you notice that the three ions have the same electronic configuration as argon?

When writing the valence configurations of ions, you must consider the charge of the ion. The charge of the ion will determine the number of electrons. A positive ion will have fewer electrons and a negative ion will have more electrons than its ground state electron configuration.

For ions, the valence equals the electrical charge. In molecules, the various atoms are assigned charge-like values such that the sum of the oxidation numbers equals the charge on the molecule. For example, in the H_2O molecule, each H has an oxidation number of +1, and the O is −2.

Example 1

Write the complete electron configuration for the chloride ion (Cl^-).

Step 1: Determine the number of electrons.

Atomic number = 17

It is a 1^- ion. Accordingly, there is one more electron than the neutral atom.

Number of electrons = 17 + 1 = 18

Step 2: Use the periodic table.

$1s^2 2s^2 2p^6 3s^2 3p^6$

Example 2

Write the complete electron configuration for the calcium ion (Ca^{2+}).

Atomic number = 20

It is a 2^+ ion. Accordingly, there are two fewer electrons than the neutral atom.

Number of electrons = 20 − 2 = 18

$1s^2 2s^2 2p^6 3s^2 3p^6$

Notice that the calcium ion and the chloride ion have the same electron configuration. They both also have the same electron configuration as the noble gas argon.

Example 3

Write the complete electron configuration for the cobalt (III) ion (Co^{3+}).

Atomic number = 27

It is a positive 3 ion. Accordingly, there are 3 fewer electrons than the neutral atom.

Number of electrons = 27 − 3 = 24

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$

Note that when a d block element loses electrons, they are lost from the 4s energy level before the 3d energy level.

Learning Activity 2.5: Electron Configuration of Ions

1. Write complete electronic configurations for the following ions:
   a. As^{3+} ____
   b. Ni^{2+} ____
   c. Fe^{2+} ____
   d. Ni^{4+} ____
   e. Zn^{2+} ____
   f. Br^- ____

continued

Learning Activity 2.5: Electron Configuration of Ions (continued)

2. Write the electronic configurations for the valence electrons of each of the following:

   a. Mg ____

   b. S^{2-} ____

   c. Cr^{3+} ____

   d. N^{3-} ____

   e. Fe^{3+} ____

3. Complete the following table:

Element Name	Electron Configuration	Valence Shell	Group Number
Helium			
Lithium			
Beryllium			
Carbon			
Nitrogen			
Oxygen			
Fluorine			
Sodium			
Magnesium			
Aluminum			
Silicon			

4. What observation can you make about the number of valence electrons and the group number of a given electron? Give an example.

Check the answer key.

Lesson Summary

In this lesson, you learned that the valence electrons are lost and gained in chemical reactions and that these electrons will always be from the s- and p-orbitals. You also related the electron configuration of an element to its valence electron configuration and to its position on the periodic table.

Assignment 2.3: Valence Electrons (10 marks)

1. Write the complete electronic configuration for the Cu+ ion. *(1 mark)*

   ____

2. Identify the elements that have the following electron configurations: *(7 marks)*

   a. $1s^2 2s^2 2p^6 3s^2 3p^3$ ____

   b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ ____

   c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ ____

   d. $1s^2 2s^2 2p^6 3s^1$ ____

   e. $1s^2 2s^2 2p^4$ ____

   f. $[Ne]3s^2 3p^5$ ____

   g. $[Ar]4s^1$ ____

3. Write the valence electron configurations for the following elements: *(2 marks)*

   a. Boron ____

   b. Sulphur ____

Notes

LESSON 5: PERIODIC TRENDS (2 HOURS)

Lesson Focus

SLO C12-2-07: Identify and account for periodic trends among the properties of elements, and relate the properties to electron configuration.
Include: atomic radii, ionic radii, ionization energy, and electronegativity

Lesson Introduction

In 1870, Dmitri Mendeleev (1834–1907) arranged the 65 elements known at that time into a periodic table. He proposed the "periodic law," which states that when the elements are arranged according to their mass, they have a regular periodic repeating of similar properties. Using the periodic law, he was able to predict the existence of several elements and their properties that were as yet undiscovered and for which he left spaces in his table. In this lesson, you will examine some periodic trends and explain them using the principles of atomic structure.

Force on an Electron

The positively charged nucleus of the atom applies a force on each negatively charged electron, holding the electrons around the atom. The force on an electron will affect its position around the atom. The force on an electron is dependent upon three factors:

1. Nuclear Charge

   The more protons that are in the nucleus, the greater will be the nuclear charge. The greater the nuclear charge, the more force there is on each electron in the atom.

2. Distance of the Electron from the Nucleus

   The further two charges are from each other, the lower their force of attraction is to each other.

3. Shielding Effect

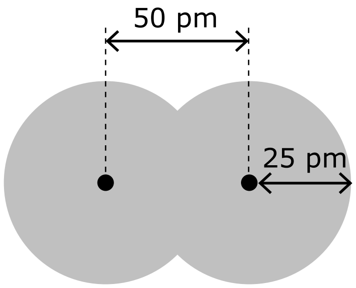
   Inner electrons tend to screen or shield the force of the nucleus from outer electrons.

The inner electrons almost cast a shadow on the outer electrons. As a result, the force of the nucleus on outer electrons is reduced. The effective nuclear charge on outer electrons becomes lower as more orbitals are added to an atom.

Atomic Radii

As you learned in previous lessons, the electron orbitals are referred to as regions of probability. As such, atoms do not have clearly defined borders. Since the border of the atom is fuzzy, chemists developed a method for measuring the size of an atom. The atomic radius is the distance between two nuclei when two like atoms are bonded together.

Scientists measure atoms by the process of **X-ray crystallography**. Very pure samples of a substance are frozen. The atoms or molecules in the resulting solid arrange very neatly, forming a characteristic structure. An X-ray that is passed through the crystalline solid will be scattered (diffracted) and the pattern that forms in a photograph of these X-rays gives information on the size of the atom, and the size and shape of the molecules.



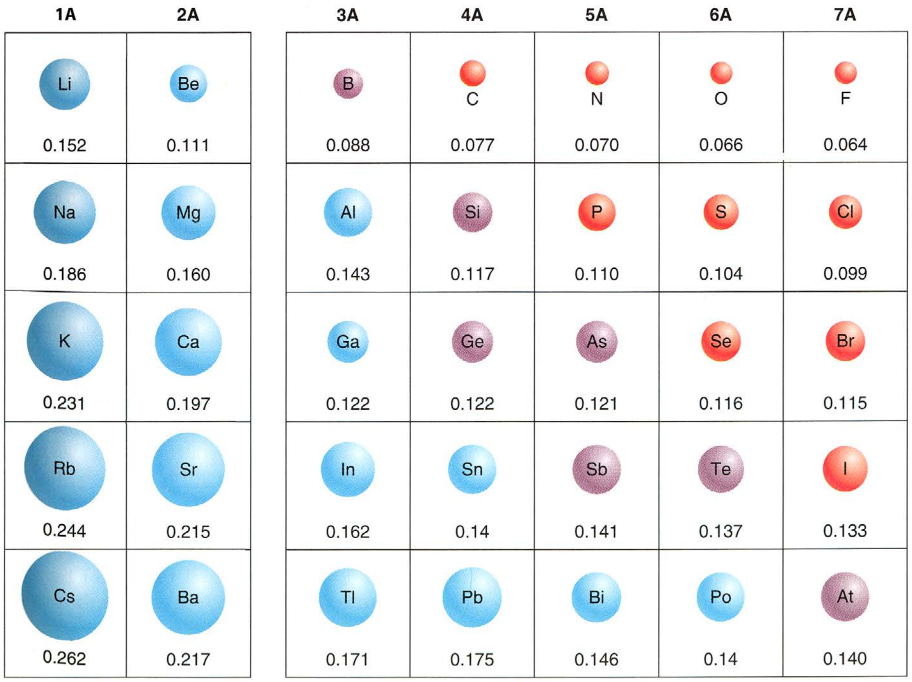
In the figure above, the distance between the two nuclei in the molecule is 50 pm (1 pm = 10^{-12}m). The atomic radius of the atom is then 25 pm. As you can see, these distances are very small.

Trends in Atomic Radii

As you move from **left to right across the periodic table**, a proton is added to each atom. For each proton added, an electron is added to the same principal energy level; however, the amount of shielding occurring on each electron remains constant. If the nuclear charge increases while the shielding remains constant, the force on each electron increases. The increased force on each electron draws the outermost electrons closer to the nucleus, and the size of the atom decreases.

As you move **down a group**, like the group 1 alkali metals, a new energy level is added for every atom in that group. With each added energy level the electrons are situated further from the nucleus, reducing the force of the nucleus on the outermost electrons. The addition of a new energy level increases the amount of shielding on the outermost electrons, resulting in a lower effective nuclear force. The lower force on each outermost electron allows the electrons to move further from the nucleus, producing a larger atomic radius.

Atomic Radii of Representative Elements (nm)



1A	2A	3A	4A	5A	6A	7A
Li 0.152	Be 0.111	B 0.088	C 0.077	N 0.070	O 0.066	F 0.064
Na 0.186	Mg 0.160	Al 0.143	Si 0.117	P 0.110	S 0.104	Cl 0.099
K 0.231	Ca 0.197	Ga 0.122	Ge 0.122	As 0.121	Se 0.116	Br 0.115
Rb 0.244	Sr 0.215	In 0.162	Sn 0.14	Sb 0.141	Te 0.137	I 0.133
Cs 0.262	Ba 0.217	Tl 0.171	Pb 0.175	Bi 0.146	Po 0.14	At 0.140

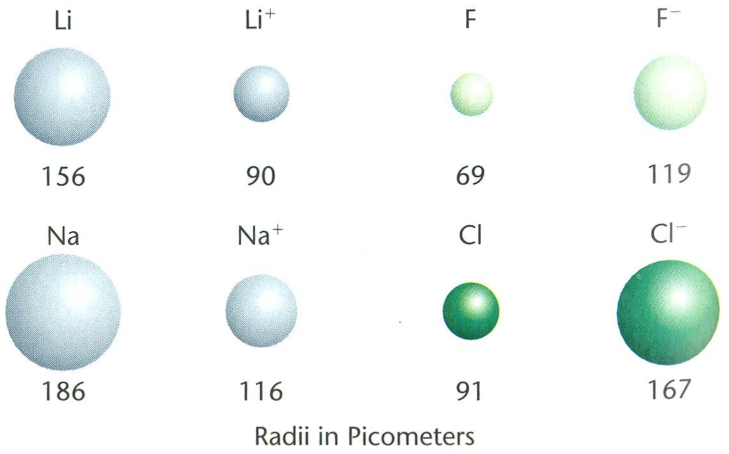
LeMay, H. Eugene, Jr., *et al. Chemistry: Connections to Our Changing World.* 2nd ed. Upper Saddle River, NJ: Prentice-Hall Inc., 2000. 175.

Note

- As you move from left to right across the periodic table, the atomic radii decrease.
- As you move down a group, the atomic radii increase.

Ionic Radii

The ratio of protons to electrons determines the force of the nucleus on each electron. In a neutral atom, there is one proton for every electron.



Phillips, John S., *et al. Chemistry: Concepts and Applications.* New York, USA: The McGraw-Hill Companies Inc., Glencoe/McGraw-Hill, 2000. 260.

For **negative ions**, one or more electrons are added to the atom. The more electrons are added to an atom, the lower the force of the nucleus on each electron and the ion becomes larger. In the diagram above, both the chloride ion and the fluoride ion are larger than the chlorine and fluorine atoms.

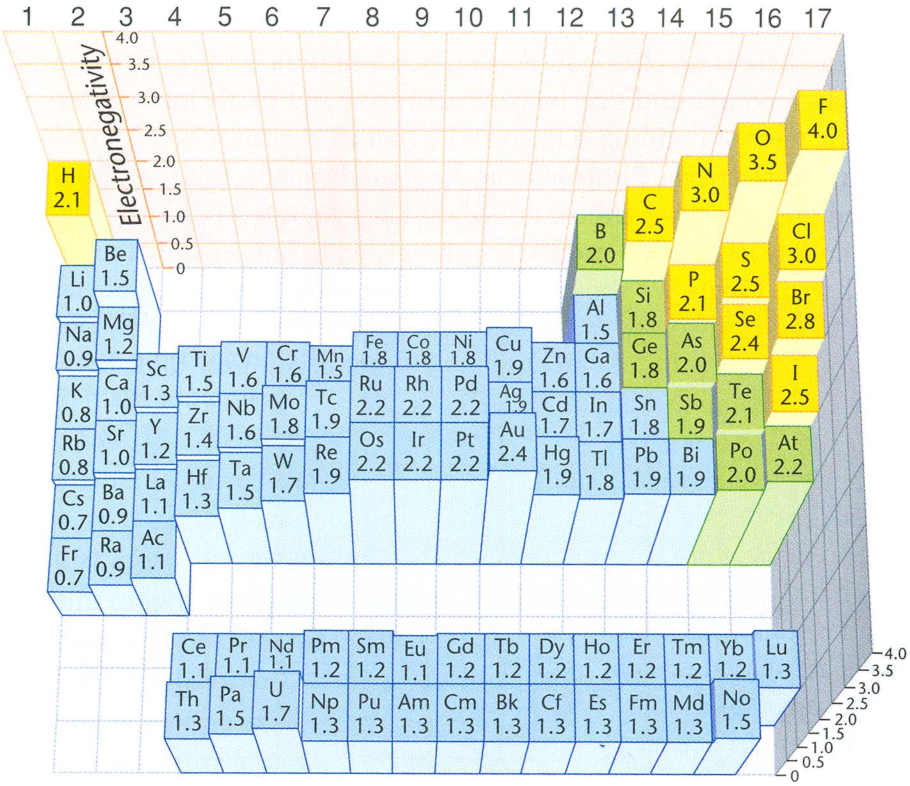
For **positive ions**, one or more electrons are removed from the atom. As a result, there is more than one proton for every electron in the ion. The force on each electron is increased, and the ion is much smaller than the atom.

Note

- As an ion becomes more negative, the ionic radius increases.
- As an ion becomes more positive, the ionic radius decreases.

Electronegativity

Electronegativity is the ability of an atom in an element to **attract** electrons when the atom is in a compound. In general, electronegativity values decrease as you go down a group and increase from left to right across the periodic table. This means that metals (on the left-hand side of the periodic table) have lower values for electronegativity than non-metals (on the right-hand side of the periodic table). For example, fluorine has a strong tendency to attract electrons when bound in a compound. It is the most electronegative element on the periodic table, with a value of 4.0. Cesium, on the other hand, is the least electronegative element on the periodic table, with a value of 0.7. As a result, cesium is more likely to lose electrons and form a positive ion rather than pull electrons to itself. Note that these patterns do **not** apply to the transition elements.



Phillips, John S., *et al. Chemistry: Concepts and Applications.* New York, USA: The McGraw-Hill Companies Inc., Glencoe/McGraw-Hill, 2000. 304.

Note

- As you move from left to right across the periodic table, the electronegativity values increase.
- As you move down a group, the electronegativity values decrease.

Ionization Energy

Ionization energy (IE) is defined as the amount of energy needed to **remove** an electron from a gaseous atom.

$$A(g) + \text{energy} \rightarrow A^+(g) + 1e^-$$

The amount of energy required depends upon the force on the electron. The greater the force on the electron, the more energy is needed to remove the electron, hence the greater the value of the ionization energy.

Periodic Trends in Ionization Energy

There are two major trends related to ionization energy and the periodic table.

Trend 1: As you move down a group, the ionization energy decreases.

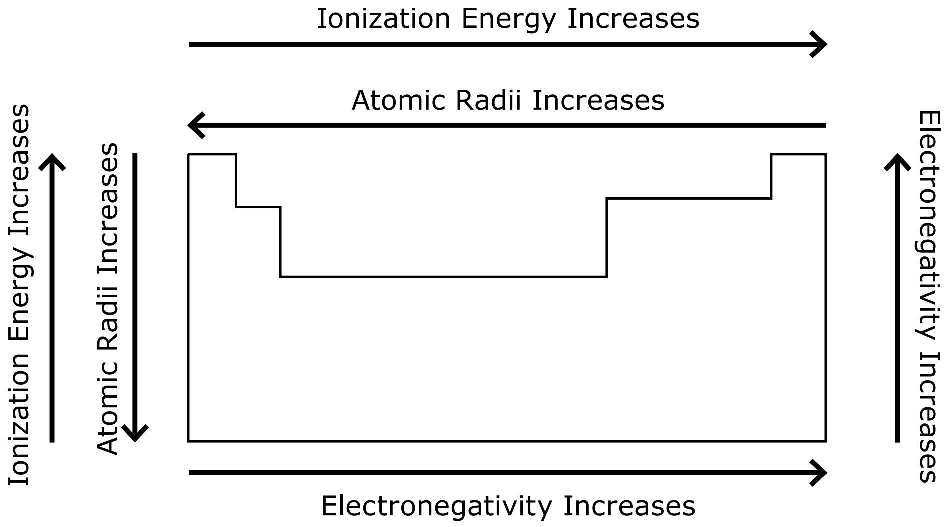
As mentioned while discussing atomic radii, as you move down a group, an energy level is added for each period. The added energy level means that there are more electrons between the outer electrons and the nucleus, which increases the shielding effect. The added energy levels also increase the size of the atoms.

The increased distance of the outer electrons from the nucleus as well as the increased shielding decrease nuclear force. The lower the force on the outer electrons, the easier they are to remove, which results in a lower IE.

Trend 2: As you move across a period from left to right, the ionization energy increases.

As you move across a period, the nuclear charge increases and, because the number of inner orbitals remains constant, the shielding remains constant. More protons in the nucleus can pull outer electrons closer and this causes the atomic radii to decrease. The decreased radius means the distance between the nucleus and the outer electrons is lower, resulting in a greater force on the outer electrons. The greater force on the electrons means it is more difficult to remove the outer electrons and it results in a higher IE.

Here is a summary diagram of the trends you have learned in this lesson.



Learning Activity 2.6: Properties Relating to Electron Configuration

4. Given the following elements and their electron configuration, which element will have the lowest ionization energy? Why?

   Element A: $1s^22s^22p^63s^23p^64s^2$

   Element B: $1s^22s^22p^63s^23p^5$

   Element C: $1s^22s^22p^63s^23p^64s^1$

   Element D: $1s^22s^22p^63s^23p^6$

5. Which of the following has the largest atomic radius and which has the smallest? Explain.
   a. nitrogen
   b. phosphorus
   c. arsenic

3. For each of the following properties, indicate whether fluorine or bromine would have the larger value.
   a. atomic radius
   b. ionization energy
   c. ionic radius

4. Arrange the following from largest to smallest ionic radii. Explain the order.

   Ne, Mg^{2+}, F^-, Na^+, O^{2-}

Check the answer key.

Lesson Summary

In this lesson, you learned about trends, or patterns, that occur on the periodic table. Below is a summary of these trends.

- As you move from left to right across the periodic table, the atomic radii decrease. As you move down a group, the atomic radii increase.
- As an ion becomes more negative, the ionic radius increases. As an ion becomes more positive, the ionic radius decreases.
- As you move from left to right across the periodic table, the electronegativity values increase. As you move down a group, the electronegativity values decrease.
- As you move across a period from left to right, the ionization energy increases. As you move down a group, the ionization energy decreases.

Assignment 2.4: Periodic Trends (14 marks)

1. For each of the following pairs of elements, indicate which would have the larger value for ionization energy. *(2 marks)*

   a. Na or K ______________________________

   b. Mg or P ______________________________

2. Arrange the following elements from smallest to largest. *(2 marks)*

   S, Cl, Al, Na ______________________________

3. Which element in each pair has the higher electronegativity value? *(4 marks)*

   a. Cl or F ______________________________

   b. C or N ______________________________

   c. Mg or Ne ______________________________

   d. As or Ca ______________________________

4. Which element in each pair has the smaller atomic radius? *(4 marks)*

   a. Na or Li ______________________________

   b. Ca or Mg ______________________________

   c. C or Ge ______________________________

   d. Se or O ______________________________

5. How does the shielding effect influence the force on an electron? *(2 marks)*

   ____

   ____

   ____

   ____

   ____

   ____

Notes

MODULE 2 SUMMARY

Congratulations, you have now completed Module 2!

In this module, you learned about the nature of light, which helped you fully understand the behaviour of an atom. Next, you studied the historical development of the quantum mechanical model of the atom. Then, you practiced writing electron configurations and valence electron configurations for the elements of the periodic table. You finished the module by identifying periodic trends on the periodic table and relating these patterns to electron configuration.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 2 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:

☐ Cover Sheet for Module 2 (found at the end of the course Introduction)
☐ Assignment 2.1: Applications of Line Spectra
☐ Assignment 2.2: Electron Configuration
☐ Assignment 2.3: Valence Electrons
☐ Assignment 2.4: Periodic Trends

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

Notes

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Notes

GRADE 12 CHEMISTRY (40S)

Module 2
Atomic Structure

Learning Activity Answer Keys

MODULE 2:
ATOMIC STRUCTURE

Learning Activity 2.1: Wavelength and Frequency

What relationship do you notice between wavelength and frequency? Refer to the previous examples to help you answer this question.

In all of the previous examples, wavelength and frequency are inversely related; that is, as wavelength increases, the frequency decreases. The longer the wavelength, the more time it will take to move past a point.

Learning Activity 2.2: Line Spectra

1. Put the following in order of increasing energy: green light, X-rays, radio, red light, ultraviolet, microwaves, blue light, and gamma rays.

   radio, microwaves, red light, green light, blue light, ultraviolet, X-rays, gamma rays

2. Describe the relationship between frequency, wavelength, and energy.

   Wavelength and frequency are inversely related; wavelength and energy are inversely related; and frequency and energy are directly related.

3. Why is light called electromagnetic radiation?

   James Clerk Maxwell determined that light is formed by electrical and magnetic waves moving perpendicularly to each other.

4. Compare and contrast line spectra and continuous spectra.

   Both are made up of various wavelengths of light. Continuous spectrum contains all wavelengths of light and line spectrum has distinct wavelengths of light.

5. Using the flame test video clip, complete the following table:

Metal Ion	Flame Colour
Potassium	**Pink**
Barium	**Green**
Calcium	**Orange/Red**
Strontium	**Red**
Lithium	**Red/Orange**
Sodium	**Yellow**
Copper	**Green**

Learning Activity 2.3: The Quantum Mechanical Model

1. Identify the contribution made to atomic theory by each of the following scientists:
    a. Bohr

    Bohr proposed that electrons are arranged around the atom in orbits. Their position depends on their energy. He also proposed an explanation for line spectra.

    b. Schrödinger

    Schrödinger developed an equation that can predict an electron's position and movement based on probability. He also developed the idea of orbitals.

2. What is the Heisenberg uncertainty principle?

    According to Heisenberg, it is impossible to know simultaneously both the momentum (or movement) and the position of an electron with certainty.

3. Complete the following table describing the s-, p-, and d-orbitals.

Orbital	s-orbital	p-orbital	d-orbital
Shape	**sphere (circle)**	**pear or dumbbell**	**varied shapes**
Number of orbitals in each principal energy level	**1 per level**	**3 per level**	**5 per level**
Found in which principal energy levels	**all**	**2nd and higher**	**3rd and higher**

4. What are orbitals?

    The s-, p-, and d-orbitals represent regions of probability with regard to where the electrons are located in three-dimensional space.

5. Compare ground state and excited state.

    When radiation is absorbed by an atom, an electron jumps from the ground state (its resting state) to a higher unstable energy level called the excited state.

6. How many orbitals (or sublevels) are found in the 4th principal energy level?

   seven

7. Complete the following sentences with the correct term.

   a. According to the Heisenberg uncertainty principle, it is impossible to know the **speed/motion** and the position of a particle at the same time.

   b. Energy is **absorbed** when an electron moves from ground state to excited state.

   c. Electron **orbitals** are areas of high probability of finding electrons.

   d. Orbitals are described in terms of size, shape, and **orientation** in space.

   e. Principal quantum numbers, also called principal energy levels, are represented by the symbol **N**.

   f. The size of the s-orbital **increases** as the principal quantum number, or energy level, increases.

Learning Activity 2.4: The Aufbau Principle

1. How many electrons in an atom can be found in the following orbital, sublevel, or principal energy level?

   a. 3d **10**

   b. n=2 **$2n^2 = 8$**

   c. 4p **6**

   d. f **14**

   e. s **2**

   f. 2p **6**

   g. 3d **10**

   h. n=5 **$2n^2 = 50$**

2. Write complete electronic configurations for the following atoms and ions:
   - a. Cu = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$**
   - b. Zn = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$**
   - c. P = **$1s^2 2s^2 2p^6 3s^2 3p^3$**
   - d. Ca = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$**
3. How many unpaired electrons are there in each of the following?
   - a. Ca = **0**
   - b. Ge = **2**
   - c. Li = **1**
   - d. O = **2**
   - e. Mn = **5**
   - f. P = **3**

Learning Activity 2.5: Electron Configuration of Ions

1. Write complete electronic configurations for the following ions:
   - a. As^{3+} = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$**
   - b. Ni^{2+} = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^8$**
   - c. Fe^{2+} = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$**
   - d. Ni^{4+} = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$**
   - e. Zn^{2+} = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^{10}$**
   - f. Br^- = **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$**
2. Write the electronic configurations for the valence electrons of each of the following:
   - a. Mg = **$3s^2$**
   - b. S^{2-} = **$3s^2 3p^6$**
   - c. Cr^{3+} = **$3s^2 3p^6 3d^3$**
   - d. N^{3-} = **$2s^2 2p^6$**
   - e. Fe^{3+} = **$3s^2 3p^6 3d^5$**

3. Complete the following table:

Element Name	Electron Configuration	Valence Shell	Group Number
Helium	**$1s^2$**	**$1s^2$**	**18**
Lithium	**$1s^22s^1$**	**$2s^1$**	**1**
Beryllium	**$1s^22s^2$**	**$2s^2$**	**2**
Carbon	**$1s^22s^22p^2$**	**$2s^22p^2$**	**14**
Nitrogen	**$1s^22s^22p^3$**	**$2s^22p^3$**	**15**
Oxygen	**$1s^22s^22p^4$**	**$2s^22p^4$**	**16**
Fluorine	**$1s^22s^22p^5$**	**$2s^22p^5$**	**17**
Sodium	**$1s^22s^22p^63s^1$**	**$3s^1$**	**1**
Magnesium	**$1s^22s^22p^63s^2$**	**$3s^2$**	**2**
Aluminum	**$1s^22s^22p^63s^23p^1$**	**$3s^23p^1$**	**13**
Silicon	**$1s^22s^22p^63s^23p^2$**	**$3s^23p^2$**	**14**

4. What observation can you make about the number of valence electrons and the group number of a given electron? Give an example.

   The number of valence electrons corresponds to the group number. For example, mangesium is in group 2 and has 2 valence electrons. Fluorine is in group 17 (or VIIB, depending on your periodic table) and has 7 valence electrons.

Learning Activity 2.6: Properties Relating to Electron Configuration

1. Given the following elements and their electron configuration, which element will have the lowest ionization energy? Why?

   Element A: $1s^22s^22p^63s^23p^64s^2$

   Element B: $1s^22s^22p^63s^23p^5$

   Element C: $1s^22s^22p^63s^23p^64s^1$

   Element D: $1s^22s^22p^63s^23p^6$

   Element C. Element C has 4 energy levels, which is more than B or D. This means that C's electrons are further from the nucleus and have greater shielding than B or D. Element C and Element A are in the same period; however, because element A is further to the right, C is a larger atom so the force on its outer electrons is lower.

2. Which of the following has the largest atomic radius and which has the smallest? Explain.
   a. nitrogen
   b. phosphorus
   c. arsenic

   They are all in the same group; therefore, nitrogen (N) has the smallest atomic radius because it has the fewest electrons, thus its shielding is less and it has the fewest number of energy levels, so the nucleus draws the electrons closer.

   Arsenic is the largest because it has the most energy levels and the most electrons. This means that arsenic has the most shielding and the lowest nuclear force on outer electrons.

3. For each of the following properties, indicate whether fluorine or bromine would have the larger value.
   a. atomic radius
   b. ionization energy
   c. ionic radius

   Bromine has the largest atomic radius because it has more electrons and more energy levels. Fluorine has the highest first ionization energy because it is the smallest atom with the least amount of shielding. This means that the force on the outer electrons in fluorine is larger than in bromine. Bromine has the larger ionic radius because both gain one electron and bromine is the larger atom.

4. Arrange the following from largest to smallest. Explain the order.

   Ne, Mg^{2+}, F^{-}, Na^{+}, O^{2-}

   All of these elements have the same number of electrons; therefore, the size will depend on the nuclear force. Mg has the most protons, followed by Na, Ne, and F. Oxygen has the least number of protons; therefore, it cannot hold the electrons as closely as the others.

GRADE 12 CHEMISTRY (40S)

Module 3
Chemical Kinetics

MODULE 3: CHEMICAL KINETICS

Introduction to Module 3

Chemistry focuses largely on chemical reactions. Some reactions, such as rusting, occur very slowly, while others, such as explosions, occur very quickly. **Kinetics** is the branch of chemistry that studies the speed, or **rate**, at which chemical reactions occur. Many chemical reactions occur in more than one step. In this study of kinetics, you will learn about the steps, or the **mechanism**, of a chemical reaction. It is important for chemists to understand the mechanism of a reaction and the factors that affect its rate. With this knowledge, chemists can control a chemical reaction in order to maximize the manufacturing of commercial products.

Note

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.

Assignments in Module 3

Note

Please remember to work on Assignment 1.9: Researching a Redox Application as you work through this course. Please submit it at the end of Module 6.

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
2	Assignment 3.1	Reaction Rate
5	Assignment 3.2	Collision Theory and Reaction Rate
6	Assignment 3.3	Potential Energy Diagrams
7	Assignment 3.4	Reaction Mechanisms
8	Assignment 3.5	Rate and Order

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.

LESSON 1: REACTION RATE (2 HOURS)

Lesson Focus

SLO C12-3-01: Formulate an operational definition of *reaction rate*. Include: examples of chemical reactions that occur at different rates

SLO C12-3-02: Identify variables used to monitor reaction rates (i.e., change per unit time, $\Delta x/\Delta t$).
Examples: pressure, temperature, pH, conductivity, colour...

Lesson Introduction

When we hear the word *rate* we often think of speed or velocity, especially in terms of rate of motion. In the chemical world, some reactions occur faster than others. Can you think of a fast reaction and a slow reaction? In this lesson, you will define reaction rate and determine methods by which the rate of a chemical reaction can be measured.

What Is Rate?

Recall that a chemical reaction is one in which a new substance is formed. It is usually expressed in the shorthand form, known as a chemical equation. The chemical equation shows the **reactants** (written on the left side of an arrow) and the **products** of the reaction (written on the right side of the arrow) as demonstrated below.

$$\text{REACTANTS} \rightarrow \text{PRODUCTS}$$

Measuring Rate

How can you measure how fast a reaction occurs? A slower reaction takes a longer period of time compared to a faster reaction. For example, an explosion is a very fast reaction because the explosive burns up in fractions of a second. Conversely, burning the wax in a candle is a slow reaction because it could take hours. However, the terms *fast* and *slow* are relative. This means there is no clear and definite way to say that one reaction is fast, while another is slow.

In Grade 10 Science, you learned that speed or velocity can also be called the rate of motion. For example, when you calculate speed, you are measuring the distance travelled over a given period of time.

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{\text{m}}{\text{s}}$$

Reaction rate is usually measured by determining the change in the amount of reactant consumed or product created over a period of time. The rate of a reaction influences how much time a reaction takes to complete. The rates of chemical reactions may not be directly determined, so another observable property must be measured. The rate of a reaction is

$$\text{rate of reaction} = \frac{\text{measured change in a property}}{\text{time for change to occur}}$$

The reaction below shows that A and B combine to form C and D and heat is given off (the reaction is exothermic).

$$A_{(s)} + B_{(s)} \rightarrow C_{(g)} + D_{(aq)} + \text{heat}$$

In this reaction, rate can be measured in terms of several different properties:

1. **Mass change over time (g of C/min):** As the reaction proceeds, reactants are used up and converted to gas. This is the typical property measured.
2. **Temperature change over time (°C/min):** As the reaction proceeds, heat is produced and the system increases in temperature.
3. **Pressure change over time (kPa/s or mmHg/s):** As the reaction proceeds, a gas is produced and the pressure of the system increases if the reaction occurs in a closed container. The faster the pressure increases, the greater the rate.
4. **Others:** Over a period of time, there are other variables that could be investigated, including a pH change, for example, a change in conductivity, a colour change.

Learning Activity 3.1: Measuring Reaction Rate

1. For each of the following ways of determining reaction rate, state the units of measurement.
   a. Temperature change over time: ______________________________
   b. Pressure change over time: ______________________________
   c. Mass change over time: ______________________________
2. Define reaction rate.

Check the answer key.

Lesson Summary

In this lesson, you learned that reaction rate can be measured by determining the change in the amount of reactant consumed or product created over a period of time. You were introduced to the various ways reaction rate can be measured: temperature change, pressure change, mass change, pH change, change in conductivity, and colour change.

Notes

LESSON 2: MEASURING REACTION RATE (3 HOURS)

Lesson Focus

SLO C12-3-03: Perform a laboratory activity to measure the average and instantaneous rates of a chemical reaction.
Include: initial reaction rate

Lesson Introduction

In the previous lesson, you learned how the rate of a chemical reaction relates to the speed at which a measurable change takes place. You were also introduced to the different units that can be used to determine the rate of a chemical reaction. In this lesson, you will now apply this knowledge to determine the average and instantaneous rates of chemical reactions.

Average Rate

Just as we can calculate the average speed of a car travelling between Dauphin and Neepawa, we can also use a similar technique to calculate the average rate of a chemical reaction.

Example 1

The distance between Dauphin and Neepawa is approximately 150 km. Jeremy drives from Neepawa to Dauphin in 1.5 hours. Calculate his average speed.

Solution:

$$\begin{aligned} \text{Speed} &= \text{distance / time} \\ &= 150 \text{ km} / 1.5 \text{ hrs} \\ &= 100 \text{ km/h} \end{aligned}$$

Average speed is calculated by using the entire distance travelled for a particular interval and dividing it by the time that the interval takes. Calculating the rate of a chemical reaction is done in the exact same way.

Example 2

Amanda is investigating the reaction between magnesium metal and hydrochloric acid. She finds that the reaction produces 45 mL of hydrogen gas in 1 minute and 15 seconds. Amanda must calculate the average rate of hydrogen gas production.

$$2HCl(aq) + Mg(s) \rightarrow H_2(g) + MgCl_2(aq)$$

Solution:

In this reaction, a measured quantity of hydrogen gas is being produced and recorded with respect to time. Both seconds and minutes are used in the question. Units must be the same, so we will convert all time units to seconds (1 minute = 60 seconds).

Amanda is ready to calculate the average rate for this reaction. Calculating the average rate of a chemical reaction is much like calculating the average speed. She will calculate the measured change in gas production with respect to time.

Average rate = measured change in quantity / change in time = Δ product / Δ time

$$= 45 \text{ mL} / 75 \text{ seconds}$$

$$= 0.60 \text{ mL/s}$$

In this example, Amanda finds that the average rate of hydrogen gas production was 0.60 mL/s.

Example 3

Amanda is to collect a set of data so that the average rate of gas production can be calculated for any particular interval of time. This will allow her to see how the rate of a chemical reaction can change as it proceeds.

Consider the same reaction but, this time, Amanda decides to collect data recorded every 5 seconds.

Time (s)	Hydrogen Gas, H_2, Produced (mL)
0	0
5	5.5
10	10.8
15	14.6
20	20.4
25	23.5
30	28.6
35	31.7
40	35.8
45	38.2
50	40.1
55	41.6
60	42.8
65	43.2
70	44.1
75	45.0

Solution:

[We will use the following notation for our calculations:

V_i = the initial volume (volume at the start of the time interval)
V_f = the final volume (volume at the end of the time interval)
t_i = the time at the start
t_f = the time at the end of the interval]

1. Calculate the average rate of gas production in mL/s for the first 20 seconds.

   Average rate $= V_f - V_i / t_f - t_i$
   $= 20.4 \text{ mL} - 0 \text{ mL} / 20 \text{ s} - 0 \text{ s}$
   $= 1.02 \text{ mL/s}$

2. Calculate the average rate of gas production in mL/s for the last 20 seconds.

   Average rate $= 45.0 \text{ mL} - 41.6 \text{ mL} / 75 \text{ s} - 55 \text{ s}$
   $= 3.4 \text{ mL}/ 20 \text{ s}$
   $= 0.17 \text{ mL/s}$

In the example, Amanda was able to come to the conclusion that the reaction between hydrochloric acid and magnesium metals slows considerably as time passes. The factors that control the rate of a chemical reaction will be explored in further detail in lesson 5.

If data are provided, then calculating the average rate for an entire set of data (or a small interval) is quite easy. It is useful to compare how the rate of a chemical reaction changes as it proceeds. For an even smaller snapshot into the data, though, one could calculate the instantaneous rate of a chemical reaction.

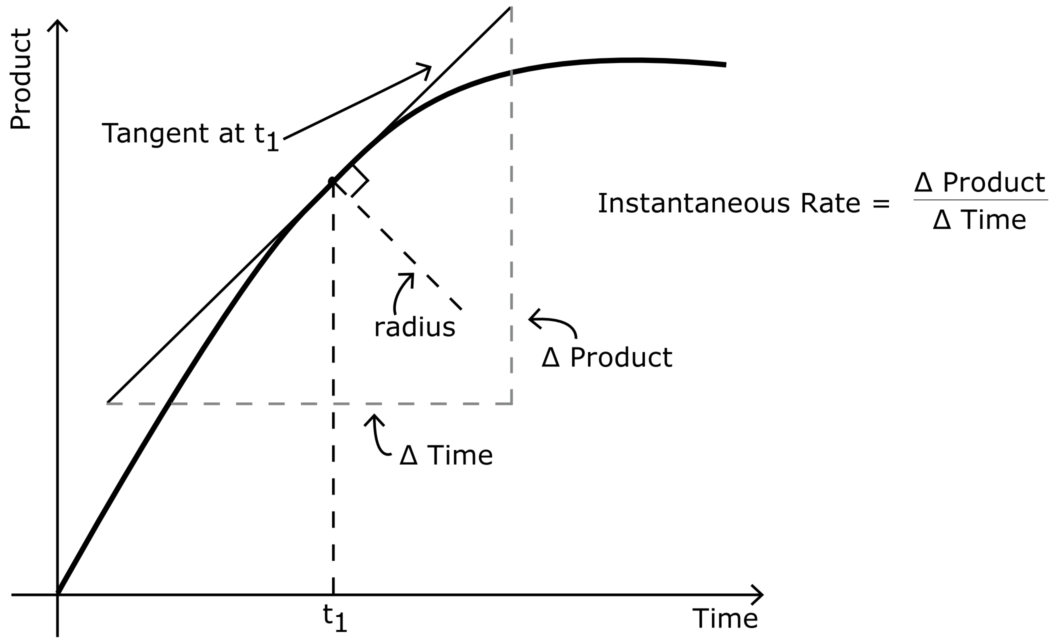
Instantaneous Rate

For accuracy, chemists often need to know exactly how fast a reaction is going at specific times, not just over a period of time. The rate of a reaction at a specific time is referred to as the **instantaneous rate**. To properly determine the instantaneous rate of a reaction, we can determine the slope of the graph at one time; however, most rate graphs are curved, not straight, so we need to calculate the slope of a **tangent** to the curve at the specific time. A tangent is a straight line that intersects a curved line through one point and is perpendicular to the radius of the curve. (See the following graph.)

The proper and most accurate method to calculate the slope of the tangent is to use calculus; however, for high school chemistry, a good approximation can be performed by hand-drawing a smooth line. The first step is to graph the experimental data and draw a smooth line joining the data points. Time is always plotted on the x-axis, and the quantity of product or reactant is plotted along the y-axis. The quantity could be mass, volume, or even pressure. The point on a graph for a specific time is found and, using a ruler, a line is drawn that intersects the point but does not cross other points on the curve. Imagining the radius of the curve at the point and placing the ruler perpendicular to that radius is the best way to proceed.

The instantaneous rate can be calculated by determining the slope of the tangent. The slope of a line is best known as **rise over run**. In the case that follows, the rise is the change in product (mass, concentration, temperature, and so on) and the run is time. In the following illustration, we see that the measured quantity of product is increasing. If the measured quantity of a reactant were graphed, it would more than likely be decreasing, as it would be used up as the reaction proceeds.

Instantaneous rate of product formation

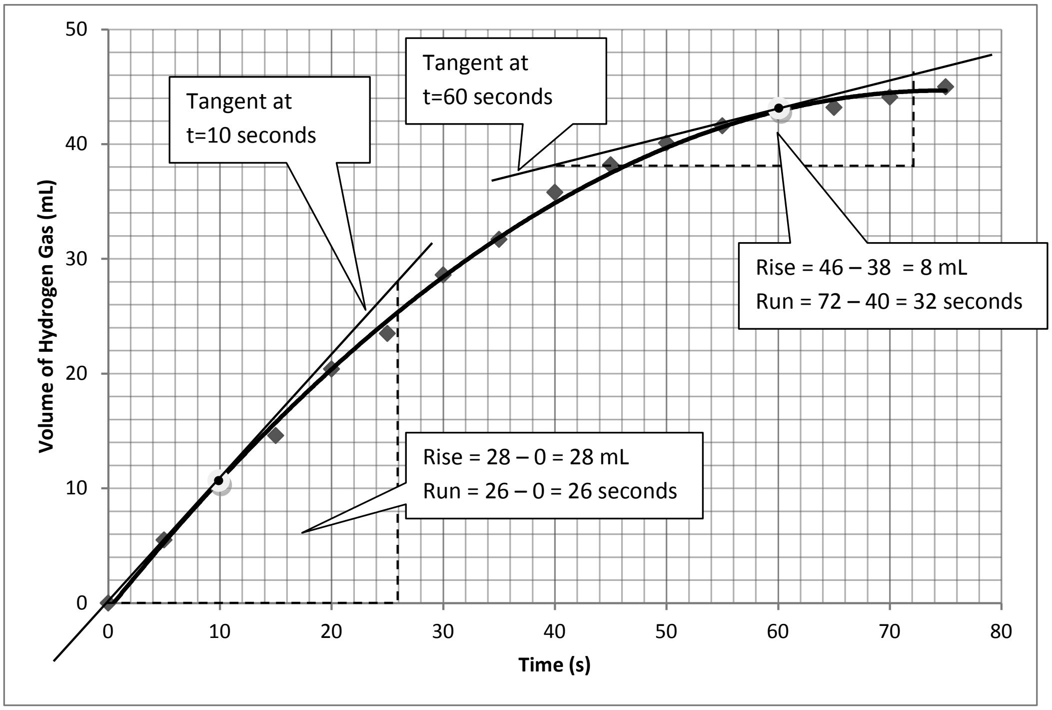


As you can see, when determining the rate for a particular instant in time for non-linear data, the tangent method can be quite useful.

Example 4

Amanda now graphed the data that she collected for the change in volume of hydrogen gas with respect to time. She joined the data lines with a smooth curve. She needed to calculate the instantaneous rate at 10 and 60 seconds; therefore, she found the points on the curve that corresponded to these times and drew two tangents.

Volume of Hydrogen Gas Produced versus Time



1. Determine the instantaneous rate at 10 seconds.
   Calculate the slope of the tangent drawn:
      Slope = rise/run = 28 mL/26 s = 1.1 mL/s.
2. Determine the instantaneous rate at 60 seconds.
   Calculate the slope of the tangent drawn:
      Slope = rise/run = 8 mL/32 s = 0.25 mL/s.

When comparing the two rates calculated above, there is a considerable decrease in the rate of the reaction. As the concentration of the reactants decreases, the reaction rate will subsequently decrease. This will be discussed in greater depth in Lesson 4.

If we compare these two instantaneous rates to the average rate values calculated in Exercise 3, we see that they are not exactly the same, even though the average values include the times for which these instantaneous values were calculated.

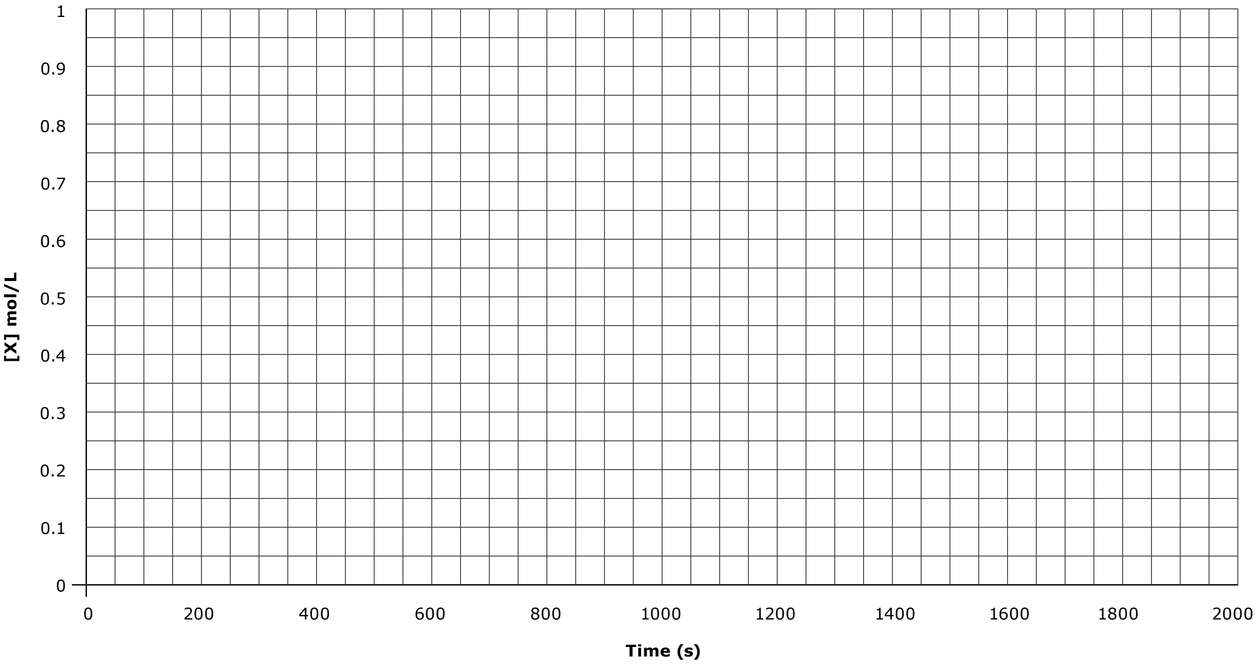
There are two things that you want to consider when drawing a tangent line.

1. Draw a long line. This will increase accuracy in determining the slope.
2. Choose points that cross at the corner of a grid. This will eliminate the need to estimate values and will provide much more accurate values.

Learning Activity 3.2: Average and Instantaneous Rates

1. Explain how average and instantaneous rates are calculated.
2. Explain the differences between average and instantaneous rates.
3. Graph the following data. Plot *time* on the x-axis and *concentration* on the y-axis.

Time (s)	[X] mol/L
0	1.00
200	0.88
400	0.78
600	0.69
800	0.61
1000	0.54
1200	0.48
1400	0.43
1600	0.38
1800	0.34
2000	0.30



continued

Learning Activity 3.2: Average and Instantaneous Rates (continued)

4. a. Explain how you would calculate the average rate of reaction for the decomposition of X over time.
   b. Calculate the average rate of the reaction in the first 200 seconds.
   c. Calculate the average rate of the reaction in the last 200 seconds.
   d. Draw a tangent at t = 400s. Calculate the instantaneous rate for this time.

Check the answer key.

Lesson Summary

In this lesson, you learned about the differences between average and instantaneous rates of chemical reactions. Data should be graphed for instantaneous rate calculations, and the slopes of tangents can be used to determine the rate for a chemical reaction at a particular instant.

Assignment 3.1: Reaction Rate (10 marks)

1. Read the following scenarios about students collecting data for a reaction rate laboratory activity and indicate which units would be used in the determination of the rate of reaction. *(1 mark each = 3 marks)*
    a. Brett collects a volume of hydrogen gas that is produced in the reaction between hydrochloric acid and magnesium metal. Data is collected every 30 seconds.
    b. Deborah is collecting data using a pressure sensor probe, which will measure the change in pressure in kPa, to determine the effect of different catalysts on the decomposition of hydrogen peroxide. Data is collected every minute for 15 minutes.
    c. Peter sets up a reaction vessel on a balance to determine the mass lost as calcium carbonate reacts with nitric acid. Data is collected every five seconds.

continued

Assignment 3.1: Reaction Rate (continued)

2. Referring to Learning Activity 3.2, Question 3, what information does the shape of the graph you constructed tell you about the reaction rate? *(1 mark)*

    ____

    ____

    ____

    ____

3. Solid calcium carbonate will react with a solution of hydrochloric acid to produce carbon dioxide gas according to the following equation:

    $$HCl_{(aq)} + CaCO_{3(s)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

    The gas is collected in a gas measuring tube and the volume of carbon dioxide gas collected is recorded at the time intervals indicated in the following chart.

Elapsed Time (s)	Volume of CO_2 (mL)
0.0	0.0
2.0	10.0
4.0	17.0
6.0	22.0
8.0	26.0
10.0	29.0
12.0	30.0
14.0	31.0
16.0	31.3
18.0	31.5

continued

a. Assuming that the reaction takes a total of 18 seconds to reach completion, what is the **average** rate of production of carbon dioxide gas measured in mL/s? *(1 mark)*

b. Calculate the rate of reaction for the first eight seconds in mL/s. *(1 mark)*

c. Calculate the rate of reaction for the last eight seconds in mL/s. *(1 mark)*

continued

Assignment 3.1: Reaction Rate (continued)

d. Explain what happens to the rate of the reaction as time passes. *(2 marks)*

e. If a graph were created for the data, state how the instantaneous rate would be calculated at time = 10 seconds. *(1 mark)*

LESSON 3: COLLISION THEORY (2 HOURS)

Lesson Focus

SLO C12-3-04: Relate the rate of formation of a product to the rate of disappearance of a reactant, given experimental rate data and reaction stoichiometry.
Include: descriptive treatment at the particulate level

SLO C12-3-06: Use the collision theory to explain the factors that affect the rate of chemical reactions.
Include: activation energy and orientation of molecules

Lesson Introduction

Scientists use theories to explain natural phenomenon. These explanations are a result of many experimental observations. Theories can be modified or discarded based on experimental evidence. In this lesson, you will study the theory that explains why some chemical reactions occur and why there are energy changes during reactions. It is important for chemists to understand the factors that affect reaction rates. By understanding those factors, the rates of some reactions may be controlled. For example, knowing what factors affect the rate of food spoilage may lead one to design methods to keep food from spoiling for longer periods of time.

Defining Reaction Rate

In Lesson 1, you learned that reaction rate is measured by determining the change in the amount of reactant consumed or product created over a period of time. In other words, reactants are being consumed as products are being created. Reaction rate can, therefore, be defined as the **rate of appearance of a product** or **rate of disappearance of a reactant**. Note that units for rate are always “amount of material/time” (such as moles/s or g/s).

Reactants disappear and products appear at rates that are proportional to the coefficients in a balanced equation. For example, consider the following reaction:

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

The balanced chemical equation shows that 2 moles of $H_2(g)$ disappear at the same rate that 2 moles of $H_2O(g)$ appear (moles are 2:2 or 1:1 ratio); however, the same is not true of the second reactant. Looking at the coefficients, you can see that the $O_2(g)$ disappears at half the rate that the $H_2O(g)$ appears (1:2 ratio).

Experimentally, the rate of change of reactant and products is proportional to the coefficients in the reaction. Accordingly, in the previous example, this ratio would be 2:1:2.

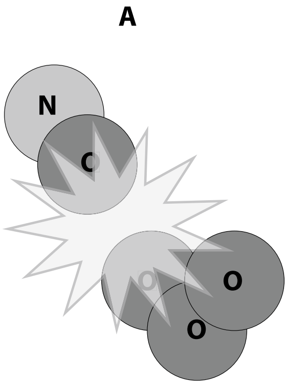
The Collision Theory

According to **collision theory**, in order for a chemical reaction to occur, the reacting particles (molecules and/or atoms) must collide with each other. If the particles do not collide, no reaction occurs.

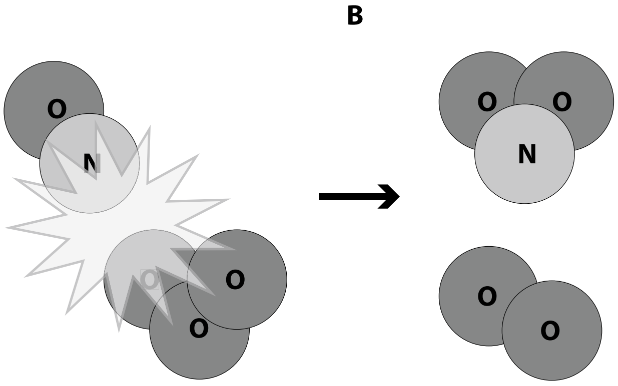
This being said, not all collisions produce a reaction. A reaction will not occur if the particles do not collide with the correct orientation (appropriate angle) and speed. This is also referred to as **reaction geometry**. In the atmosphere, for example, ozone is converted to oxygen gas and nitrogen dioxide by reacting with nitrogen monoxide, according to the following reaction:

$$NO_{(g)} + O_{3(g)} \rightarrow NO_{2(g)} + O_{2(g)}$$

If the oxygen atoms collide in orientation (diagram A), no reaction occurs. If, however, the nitrogen atom collides with the oxygen atoms (diagram B), a reaction occurs.



No reaction occurs.



A reaction occurs.

Activation Energy

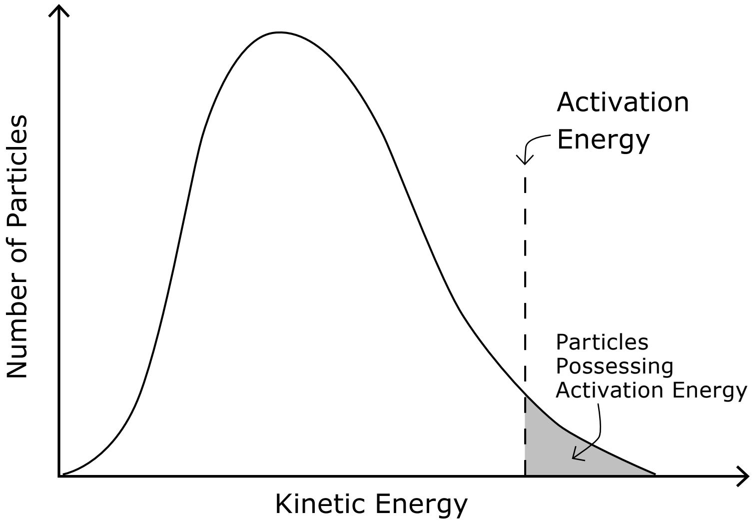
Chemical reactions involve the making and breaking of bonds, which requires energy. Not only must particles collide with the correct orientation or reaction geometry, they must also collide with enough *kinetic energy*. If colliding particles do not have sufficient velocity (kinetic energy), they will not produce a reaction.

The minimum amount of kinetic energy required for colliding particles to produce a chemical reaction is called the **activation energy** (**E_A**) of that reaction.

The greater the activation energy, the slower the reaction rate and the more time the reaction takes. For example, hydrogen and oxygen can be kept in the same container at room temperature without a reaction occurring. Even though the molecules collide, they do not possess the activation energy. However, if the mixture is heated to 800°C, or a flame or spark is introduced, an explosive exothermic reaction occurs. The heat, flame, or spark provides enough energy for the particles to reach the activation energy. They are able to strike each other with enough energy to form new hydrogen-oxygen bonds.

Kinetic Energy Distribution

You may recall from Grade 11 Chemistry that **James Clerk Maxwell** and **Ludwig Boltzman** performed experiments in 1866 which demonstrated that not all particles at a specific temperature had exactly the same velocities. In fact, some moved quite slowly while others moved quite quickly. Most, however, fell somewhere in between. Maxwell and Boltzman also found that the velocities of these particles were directly related to the amount of kinetic energy (KE or E_K) they possessed. They plotted a curve to reflect the amounts of kinetic energy possessed by the particles. This curve is known as the **Maxwell-Boltzman curve** or **kinetic energy distribution curve**.



The area under the curve represents the number of particles at a given kinetic energy. The area under the curve *to the right of the activation energy* represents the number of particles with sufficient kinetic energy to produce a collision capable of resulting in a reaction. Each reaction has its own specific activation energy. As the activation energy decreases, more particles are likely to possess enough kinetic energy to produce a reaction.

Learning Activity 3.3: The Collision Theory

1. Complete the following statements.

   a. The collision theory states that ________________________________ must collide in order to react.

   b. An effective collision leads to the formation of _____________________________.

   c. In a(n) ________________________________, particles must be oriented in a favourable position that allows the bonds of reactants to break.

   d. The energy required to break bonds between reactants comes from the ________________________________ of reactant particles.

   e. When particles collide with energy equal to the activation energy, existing ________________________________ are disrupted.

Check the answer key.

Lesson Summary

In this lesson, you learned that molecules must collide in a specific manner in order to react. If these conditions are not met, reactants will not break apart and form products. These conditions are described by the collision theory. The minimum amount of energy required for colliding particles to produce a chemical reaction is called the activation energy (E_A) of that reaction. In the next lesson, you will learn more about how energy changes during a chemical reaction can be represented graphically.

Notes

LESSON 4: REACTION RATE AND COLLISION THEORY (3 HOURS)

Lesson Focus

SLO C12-3-08: Describe qualitatively the relationship between factors that affect the rate of chemical reactions and the relative rate of a reaction, using the collision theory.

Lesson Introduction

All chemical reactions result in the breaking of some bonds (a process that needs energy) and the making of new ones (a process that releases energy). Obviously, some bonds between reactants must be broken before new ones can be made. Activation energy is involved in breaking some of the original bonds. Where collisions are relatively gentle, there isn't enough energy available to start the bond-breaking process, and the particles don't react. In this lesson, you will use the collision theory to explain which factors affect the rate of chemical reactions.

The Nature of Reactants

When considering reaction rate, the nature of reactants is definitely an important factor. Chemical reactions involve the breaking of bonds between atoms in the reactants and the reforming of bonds to make the products. The breaking of bonds requires energy while the forming of bonds releases energy. Therefore, it makes sense that the strength and number of bonds to be broken affect the rate of a reaction. The following are three ways that the nature of reactants generally affects the reaction rate:

1. The weaker the bonds to be broken, the faster is the reaction.

   In general, covalent bonds are much stronger than ionic bonds. As a result, reactants that require the breaking of covalent bonds will be slower than those involving ionic bonds. Ionic compounds in aqueous solutions are already separated into ions, so reactions in solution are usually very fast.

2. When comparing reactions with similar bonds, the greater the number of bonds that must be broken, the slower the reaction.

   Think about taking apart an object you want to fix. The more parts that you need to disassemble, the longer it generally takes.

   For example, let’s compare two reactions involving the breaking of covalent bonds. The first reaction takes place in the atmosphere and is relatively quick because it only involves the breaking of one O-O bond and 2 N-O bonds. The second reaction is the burning of octane (found in gasoline). This reaction is very slow at room temperature. This allows gasoline to be stored in the tanks of cars or in containers, because there are a large number of bonds that must be broken: 7 C-C bonds, 18 C-H bonds, and 25 O-O bonds.

   $$2\ NO_{(g)} + O_{2(g)} \rightarrow 2\ NO_{2(g)}$$

   and

   $$2\ C_8H_{18(g)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(g)}$$

3. If the reactions have similar bonds and similar numbers of bonds, the state of the reactants is important.

   Reactions with aqueous reactants tend to have the highest reaction rates because there are usually fewer bonds to break. Reactions between gaseous reactants are also fairly quick, because of the constant motion of the gas molecules and the greater likelihood of a collision occurring between the reactant particles. Reactions between solids are usually the slowest because the forces of attraction between the particles in a solid must be overcome before they can react.

Temperature Changes

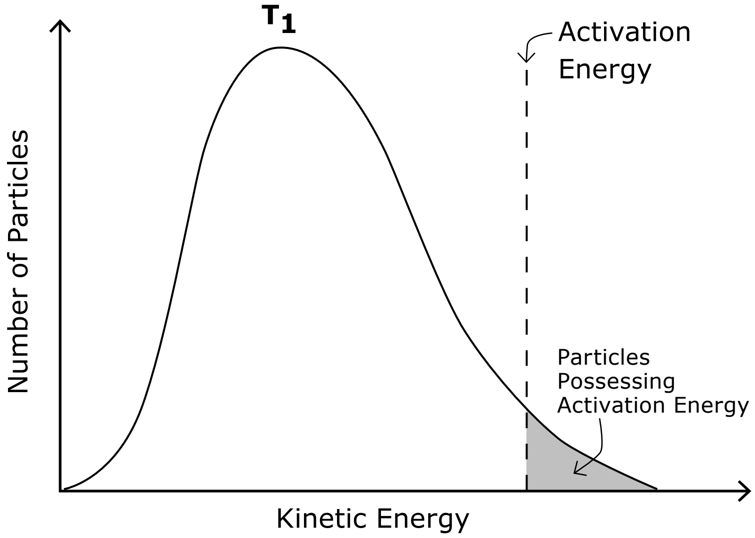
Temperature change has a direct effect on reaction rate. According to the collision theory, the rate of a reaction is determined by the frequency or number of successful or effective collisions. An effective or successful collision is a collision between reactant particles with enough energy and the correct orientation to produce a product. At any given temperature, only a fixed number of particles possess enough energy to exceed the activation energy and produce an effective collision.

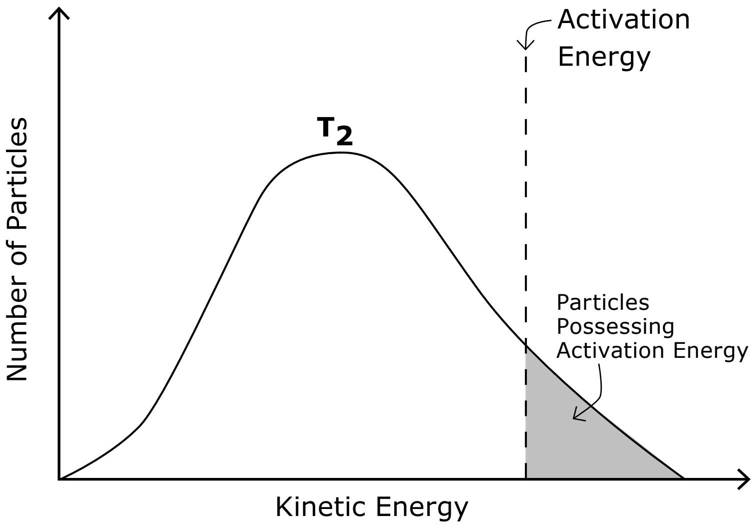
In the video found at www.dlt.ncssm.edu/core/Chapter15-Kinetics/glowsticks.htm, you can see how temperature affects the brightness of a glow stick. This is also an easy experiment you can recreate at home.

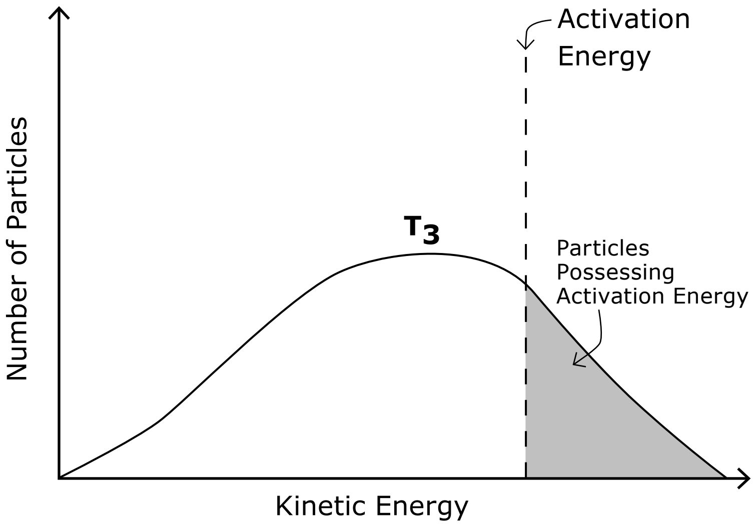
Temperature and Kinetic Energy

As the temperature increases, the **average kinetic energy** of the particles increases and the number of particles possessing the required activation energy increases. Since the number of particles with activation energy increases, the number of *successful collisions* will also increase. Increasing kinetic energy also increases the speed of the particles. If the particles are moving faster, the collisions will be more frequent.

Lowering the temperature reduces the average kinetic energy of the particles. This reduces the speed of the reactant particles and the number of reactant particles with sufficient kinetic energy. As a result, the frequency of effective collisions is decreased and the reaction rate decreases. In the graphs that follow, you can observe how the number of particles possessing the required kinetic energy increases considerably with increased temperature. Notice that changing the temperature does **not** affect the activation energy.

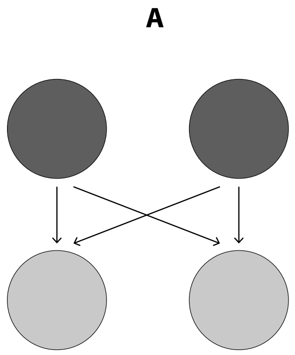


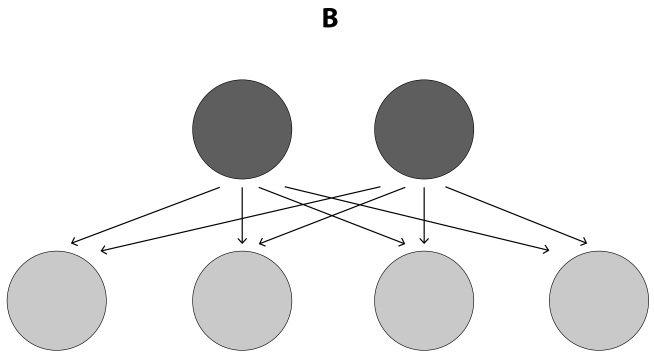




Concentration Changes

Changing the concentration of one or more reactant can affect reaction rate. Increasing the concentration of reactants increases the total number of particles in a container. If the number of particles in a container increases, the number of collisions increases and so does the number of particles with the required activation energy. More particles having the necessary activation energy correlates to an increase in the frequency of effective collisions. The following diagram illustrates the effect of doubling the concentration of just one of the reactants.





In Diagram A, with two particles of each, there exist four possible collisions that could produce a reaction.

In Diagram B, by doubling the number of light grey particles, the number of possible collisions increases to eight. This increases collision frequency.

Here is an analogy to consider: During rush hour, there are more accidents because there are more vehicles on the road at one time. The frequency of collisions increases because the concentration of vehicles is larger and the space between each vehicle is less, thus increasing the odds of a collision occurring.

Catalysts

A catalyst is a substance that accelerates or initiates a reaction without itself being permanently changed. Since the catalyst is not used up, only small amounts are required. The catalyst lowers the activation energy of a chemical reaction, which means that particles with lower kinetic energies can cause a reaction. Essentially, lowering the activation energy for a reaction allows more reactant particles to participate. Many different types of substances can act as catalysts, including organic catalysts known as **enzymes**.

Enzymes

Enzymes are known as biological catalysts. Most biological reactions, including the digestion of food and the clotting of blood, are controlled by enzymes. Enzymes are capable of increasing the rate of biological reactions by more than one million times. For example, simple sugars in food can be digested in the body in a matter of minutes.

An image showing the basic function of an enzyme as a catalyst is available at www.dlt.ncssm.edu/core/Chapter15-Kinetics/Chapter15-Animations/enzyme.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm#kinetics
2. scrolling down the page until you find the animation titled "**EnzymeCatalyst.html**" or "**EnzymeCatalyst.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Pressure Changes

Changing the pressure on a system usually only affects the reaction rates of gaseous reactions. In Grade 11 Chemistry, you learned that pressure is the force of particles upon the walls of their container. If the number of gaseous particles in a container increases without changing volume, the pressure increases.

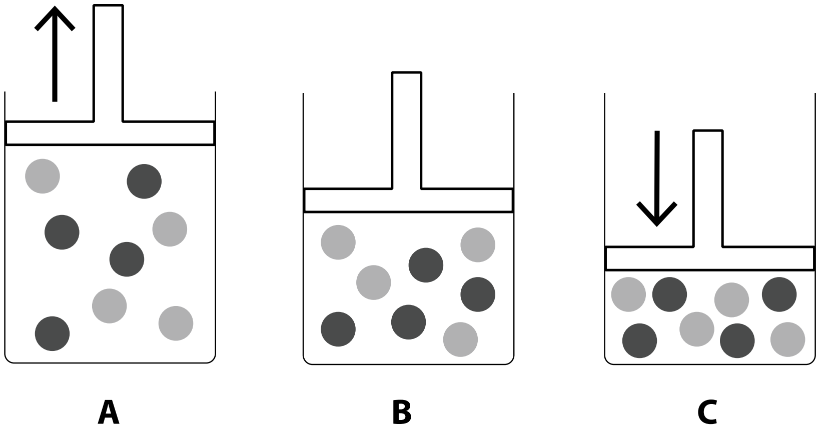
There are three ways to change pressure.

- Add more product and/or gaseous reactant particles to the container.
- Increase the volume of the container.
- Decrease the volume of the container.

If pressure is increased by adding more gaseous reactant particles, the reactant concentration increases, which causes an increased rate. A higher concentration of particles will usually result in more effective collisions. If the pressure is reduced by removing reactant, the rate decreases due to decreased reactant concentration.

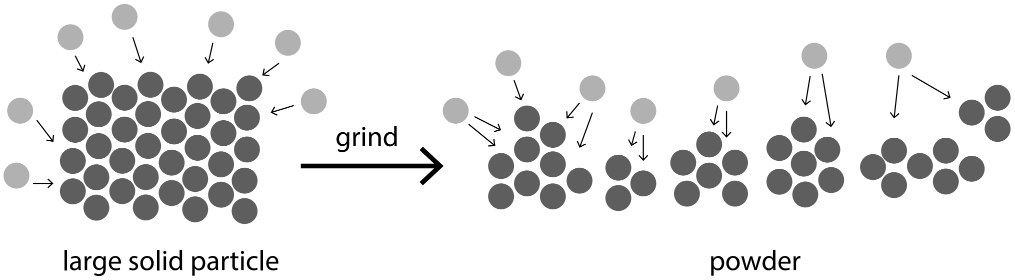
Scenario 1: The volume of container B is increased without changing the number of particles. The result looks like container A in the following diagram, where the concentration of the reactants decreases. As the spaces between the particles increase, the chances of a collision occurring decrease. Decreasing the concentration of reactants decreases the reaction rate.

Scenario 2: The volume of cylinder B is decreased without changing the number of particles in the container. The result looks like container C in the following diagram, where the concentration of the reactants increases. Since the spaces between the particles decrease, the chances of a collision occurring increase. If the concentration of the reactants increases, the reaction rate increases.



Surface Area

The size of the reactant particles (this time “particles” does not refer to molecules or atoms, but the visible “chunks” of a substance), or the surface area in contact, sometimes affects the reaction rate. Increasing the surface area of the reactants by crushing or grinding increases the number of reactant particles that come in contact with each other. The rate of reaction increases when the surface area in contact also increases.



To demonstrate the effect of surface area on the rate of reaction, use a video hosting site such as www.youtube.com/ to search for a video showing the decomposition of H_2O_2 using potatoes.

The potato cut into pieces has a greater surface area than a single cube of potato. Notice how the increased surface area affects the reaction rate.

Summary of Factors Affecting Reaction Rate

- Increasing temperature increases reaction rate by increasing the speed and number of collisions of reacting particles, as well as increasing the number of particles with the required activation energy.
- Increasing the concentration of reactants increases reaction rate by increasing the number of particles with the required activation energy and by increasing the frequency of collisions.

- Pressure changes only affect reactions containing gaseous reactants. Increasing pressure decreases the space between the particles, thus increasing the number of collisions.
- A catalyst is a substance that speeds up a reaction by reducing the required activation energy. A reduction in the activation energy increases the number of particles that can produce a successful collision, but it does not change the number of collisions between reacting particles.
- Grinding or powdering a solid reactant will increase reaction rates because of the increased surface area in contact with other reactants.

Learning Activity 3.4: Factors Affecting Reaction Rate

1. In general, what effect does an increase in the concentration of the reactants have on the rate of the reaction? Explain the effect using the collision theory.
2. How do changes in each of the following factors affect the rate of a chemical reaction?
    a. Temperature
    b. Particle size
    c. Pressure
3. List two ways to increase the rate of the following reaction:

$$Ca_{(s)} + 2\,H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

For each, use your knowledge of collision theory to explain why the rate increases.

Check the answer key.

Lesson Summary

In this lesson, you learned that the number and strength of the bonds which must be broken affect the rate of a reaction. You also learned that reaction rate is affected by temperature, reactant concentration, pressure changes (for reactions containing gaseous reactants), the addition of a catalyst, and the surface area in contact with other reactants.

LESSON 5: FACTORS AFFECTING REACTION RATE (2 HOURS)

Lesson Focus

SLO C12-3-05: Perform a laboratory activity to identify factors that affect the rate of a chemical reaction.
Include: nature of reactants, surface area, concentration, pressure, volume, temperature, and presence of a catalyst

Lesson Introduction

One of the best ways to investigate topics in any science course, particularly chemistry, is to perform a laboratory activity or to experience hands-on activities. This lesson will offer the next best thing. You will be given a laboratory activity outline with instructions. The data that was collected from the laboratory activity will be provided to you as well. You will then use this data to answer questions and form a conclusion.

Learning Activity 3.5: Reaction Rate Laboratory Activity

Investigate the effect of temperature and surface area on the reaction rate between effervescent antacid tablets and water.

If you have access to an effervescent antacid tablet and laboratory equipment, this is a very easy activity to do yourself. If you can, collect your own data by following the instructions below.

Goal/Purpose:

The goal of this activity is to

- determine the effect of temperature on the reaction rate between effervescent antacid tablets and water
- determine the effect of surface area on the reaction rate between effervescent antacid tablets and water

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

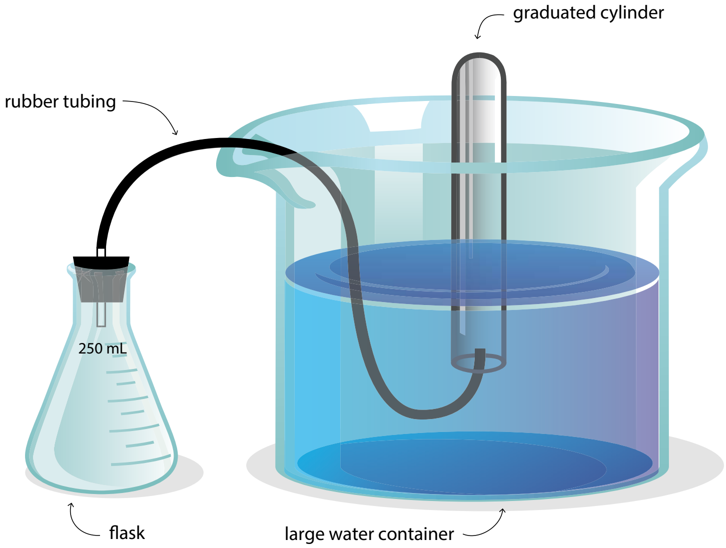
Materials Needed:

- thermometer
- 100 ml of water for each trial
    - 3 x 100 ml at 20°C (room temperature)
    - 100 ml at 30°C
- four effervescent antacid tablets
- spoon
- stopwatch
- 100 ml graduated cylinder
- 250 ml Erlenmeyer flask
- rubber tubing
- plastic bucket

Procedure:

Part A: Effect of water temperature on the reaction rate.

1. Set up the gas collection apparatus as the diagram illustrates. The upside-down graduated cylinder needs to be full of water for each trial.



2. Measure 100 ml of water at room temperature (20°C) and add this to the Erlenmeyer flask.

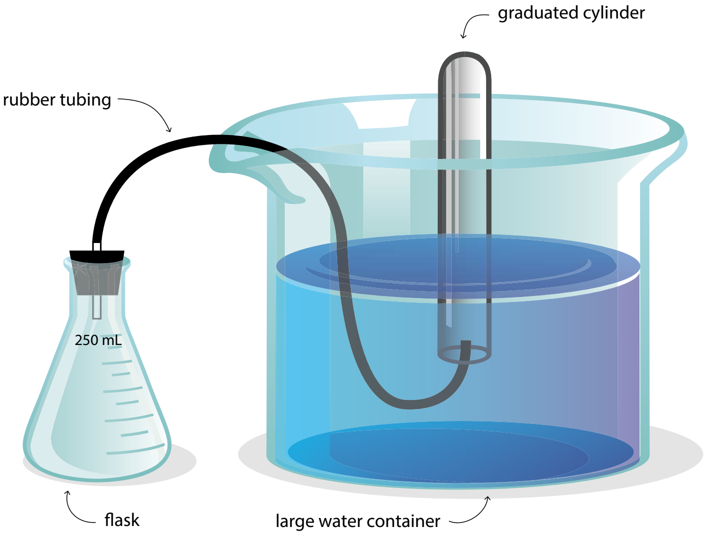
continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

3. Add one entire antacid tablet to the Erlenmeyer flask and put the rubber stopper with the rubber tubing back on the flask. Start your stopwatch.
4. Read the volume of gas collected in the graduated cylinder every 10 seconds, and record this in the data table that follows.
5. Continue to collect data until the reaction has stopped (gas production has stopped).
6. Repeat steps 1 through 5 but this time, heat the water to 30°C.

Part B: Effect of surface area on the rate of a reaction.

1. Set up the gas collection apparatus as the diagram illustrates.



2. Measure 100 ml of water at room temperature (20°C) and add this to the Erlenmeyer flask.
3. Add one entire antacid tablet to the Erlenmeyer flask and put the rubber stopper with the rubber tubing back on the flask. Start your stop watch.
4. Read the volume of gas collected in the graduated cylinder every 10 seconds and record this in the data table that follows.
5. Continue to collect data until the reaction has stopped (gas production has stopped).
6. Repeat steps 1 through 5 but this time, using a spoon (or a mortar and pestle if you have access to it), crush the tablet into a fine powder.

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Data Collection:

The following is data that has been collected from this laboratory activity. If you have the ability to collect your own data, please do so. If you do not, use this data to answer the conclusion questions as part of learning activity 3.5.

Part A: 20°C water with one entire tablet

Time (s)	Volume of Gas Collected (mL)
10	0
20	1.5
30	3.5
40	6.0
50	8.5
60	10.3
70	12.5
80	15.0
90	17.1
100	20.0
110	22.3
120	24.1
130	25.6
140	26.2
150	27.0
160	27.4
170	27.8
180	27.9

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Part A: 30ºC water with one entire tablet

Time (*s*)	Volume of Gas Collected (mL)
10	0
20	3.0
30	7.0
40	12.0
50	17.0
60	20.6
70	25.0
80	30.0
90	34.2
100	40.0
110	44.6
120	48.2
130	51.2
140	52.4
150	54.0
160	54.8
170	55.6
180	55.8

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Part B: 20°C water with one entire tablet

Time (s)	Volume of Gas Collected (mL)
10	0
20	1.5
30	3.5
40	6.0
50	8.5
60	10.3
70	12.5
80	15.0
90	17.1
100	20.0
110	22.3
120	24.1
130	25.6
140	26.2
150	27.0
160	27.4
170	27.8
180	27.9

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Part B: 20°C water with one entire tablet crushed into a fine powder

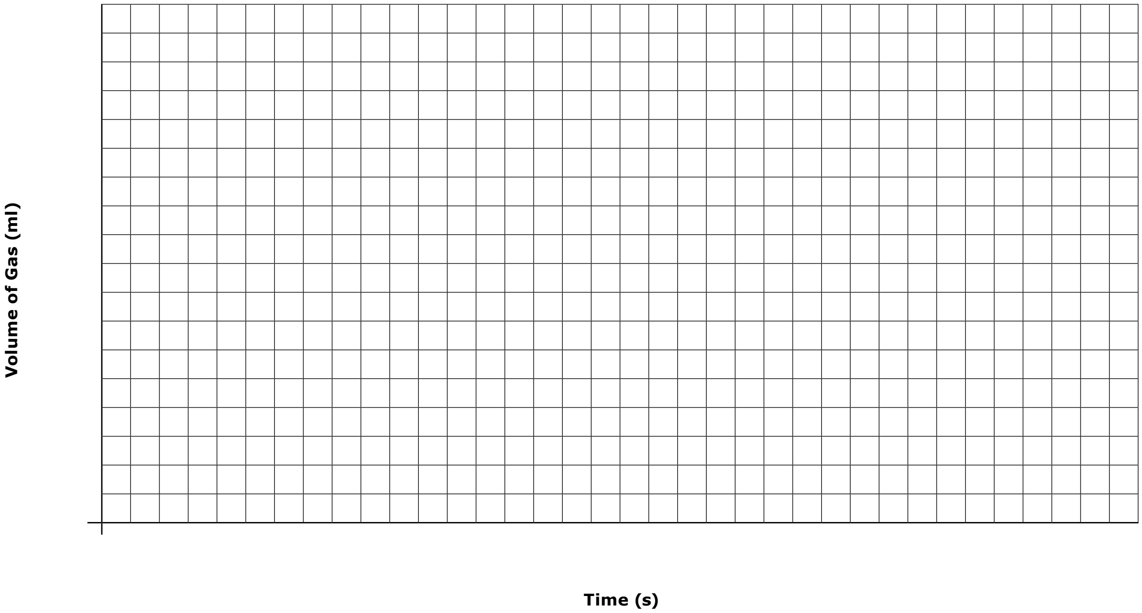
Time (s)	Volume of Gas Collected (mL)
10	0
20	4.5
30	10.5
40	18.0
50	25.5
60	30.9
70	37.5
80	45.0
90	51.3
100	60.0
110	66.9
120	72.3
130	76.8
140	78.6
150	81.0
160	82.2
170	83.4
180	83.7

1. Create a graph for each set of data provided **or** use the data you collected to create a graph.

continued

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Title ______________________________________



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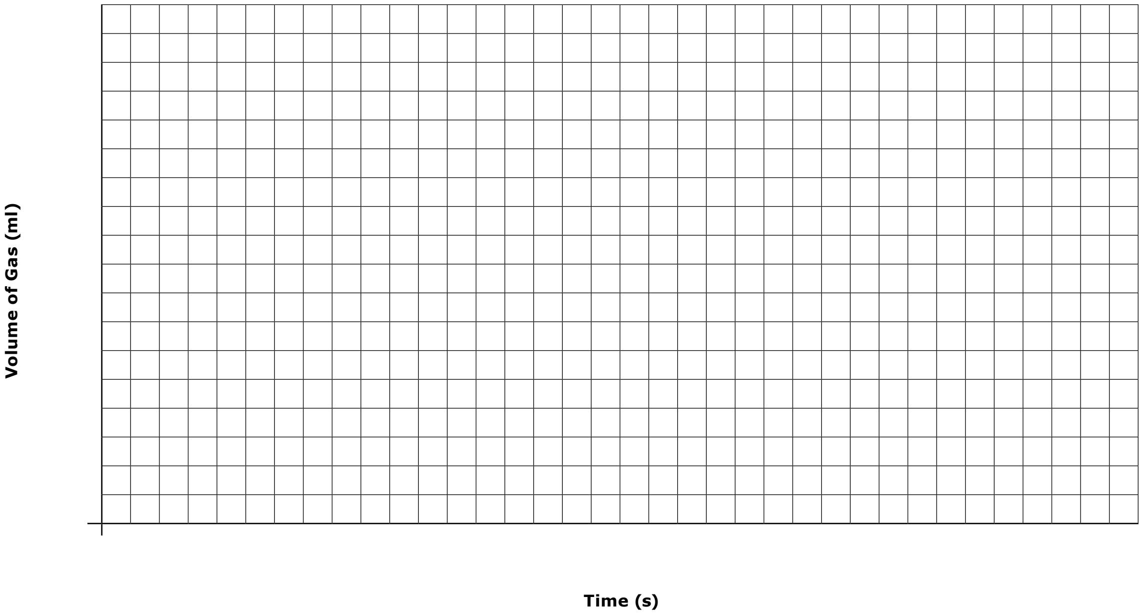

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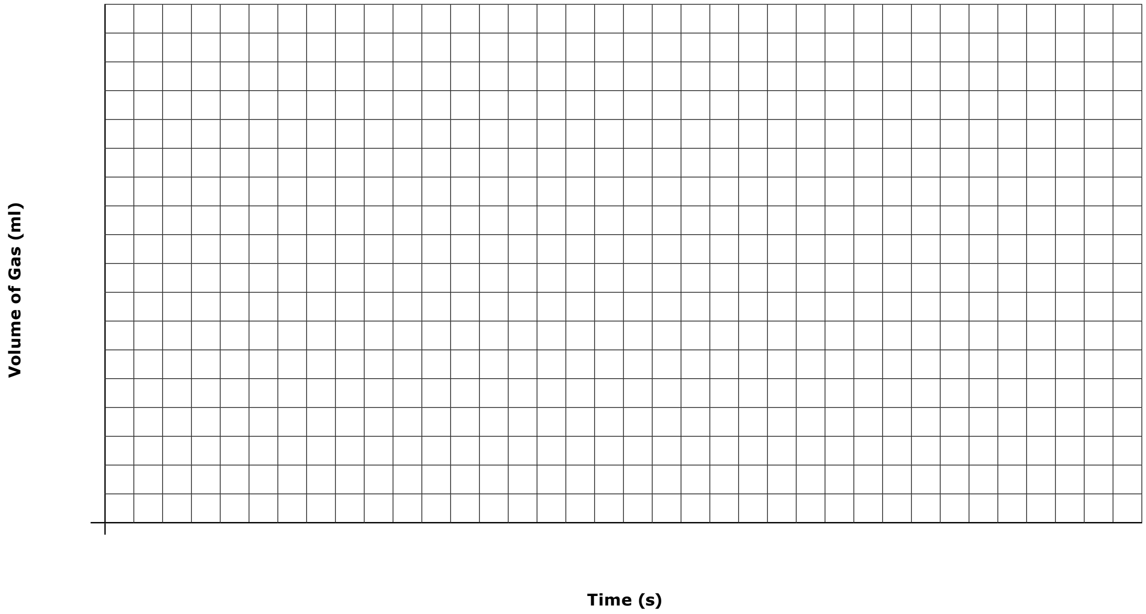
Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

Title ____



2. Work with Part A data to answer the following questions:
   a. What is the average rate for Trial 1 (20°C and one entire tablet)?
   b. What is the average rate for Trial 2 (30°C and one entire tablet)?
   c. What effect did the increase in temperature by 10°C have on the amount of gas collected?
   d. What effect did the increase in temperature by 10°C have on the reaction rate of gas collected?
   e. Using the collision theory, explain your results.

Learning Activity 3.5: Reaction Rate Laboratory Activity (continued)

3. Work with part B data for question 3.
    a. What is the average rate for trial 3 (20°C and one entire tablet)?
    b. What is the average rate for trial 4 (20°C and one entire tablet crushed into a fine powder)?
    c. What factor was manipulated by crushing the tablet into a fine powder?
    d. What effect did crushing the tablet into a fine powder have on the amount of gas collected?
    e. What effect did crushing the tablet into a fine powder have on the reaction rate of gas collected?
    f. Using the collision theory, explain your results.

Check the answer key.

Lesson Summary

This lesson offered an opportunity to collect data of your own, or to use data that had been collected, in order to investigate the effect of temperature and surface area on the reaction rate between effervescent tablets and water. If you took the opportunity to do this yourself, your data may have been different but the end results would be similar. An increase in temperature will increase the reaction rate. An increase in temperature of 10°C will approximately double the reaction rate. By crushing the tablet, the surface area is increased and this also increases the reaction rate. In collecting laboratory data, we are looking for trends in the relationships between two variables, such as temperature and reaction rate or surface area and reaction rate.

Assignment 3.2: Collision Theory and Reaction Rate (10 marks)

1. List four ways to increase the rate of the following reaction. For each, use your knowledge of the collision theory to explain why the rate increases. *(8 marks)*

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

continued

Assignment 3.2: Collision Theory and Reaction Rate (continued)

2. Explain how chewing an antacid tablet before swallowing it will affect the speed at which it will take effect. *(2 marks)*

LESSON 6: POTENTIAL ENERGY DIAGRAMS (3 HOURS)

Lesson Focus

SLO C12-3-07: Draw potential energy diagrams for endothermic and exothermic reactions.
Include: relative rates, effect of a catalyst, and heat of reaction (enthalpy change)

Lesson Introduction

The reaction of 227 grams of nitroglycerine releases 1427 kilojoules of heat, while converting 1 mole of a liquid into 29 moles of four different gases. While this reaction gives off heat, not all chemical reactions do. In this lesson, you will learn how to identify reactions that are heat producing and heat absorbing by their heat of reaction and by using potential energy diagrams.

Heat of Reaction

The rate of a reaction is determined by the amount of energy colliding particles must have in order to form products. This is the amount of energy required to produce the **activated complex**, an unstable particle that occurs at the **transition state**. You can see this on the diagram that follows. The activated complex has a maximum amount of potential energy, but exists for only a very small instant in time.

Enthalpy (**H**) is the heat content or total energy possessed by the particles in a system. The energy released or absorbed by a reaction is called the **change in enthalpy**, **ΔH**, or **heat of reaction**:

$$\Delta H = H_{PRODUCTS} - H_{REACTANTS}$$

ΔH is the change in enthalpy or energy that occurs during a reaction measured in **joules** (**J**).

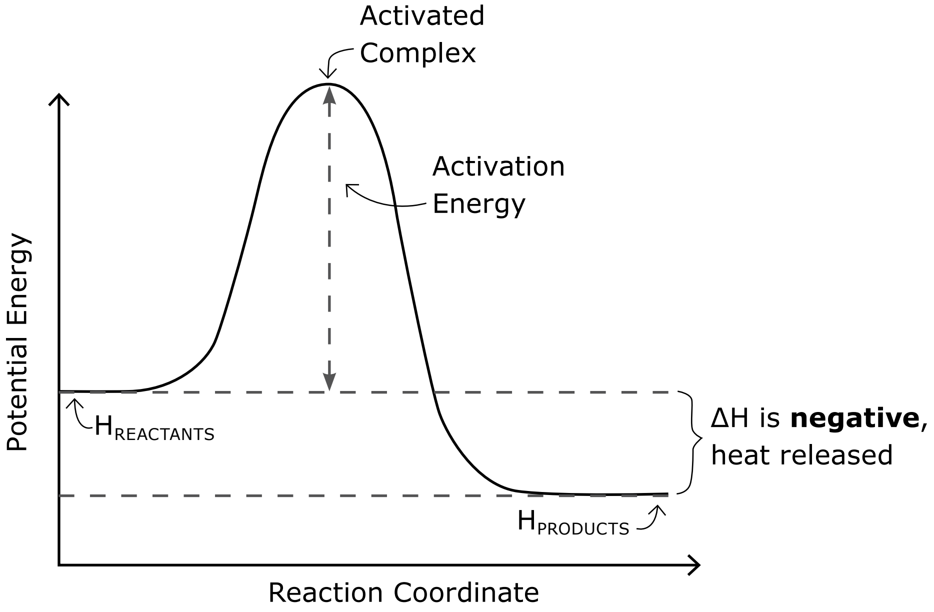
If ΔH is *negative*, the products have less enthalpy (energy) than the reactants. This indicates that heat flows out of the system. This type of reaction gives off heat and the reaction vessel feels warmer. This is called an **exothermic reaction**.

If ΔH is *positive*, heat is absorbed, or flows into the system, because the products have more enthalpy than the reactants. The reaction vessel feels cooler as energy is absorbed from the surroundings. This type of reaction is called an **endothermic reaction**.

Potential Energy Diagrams

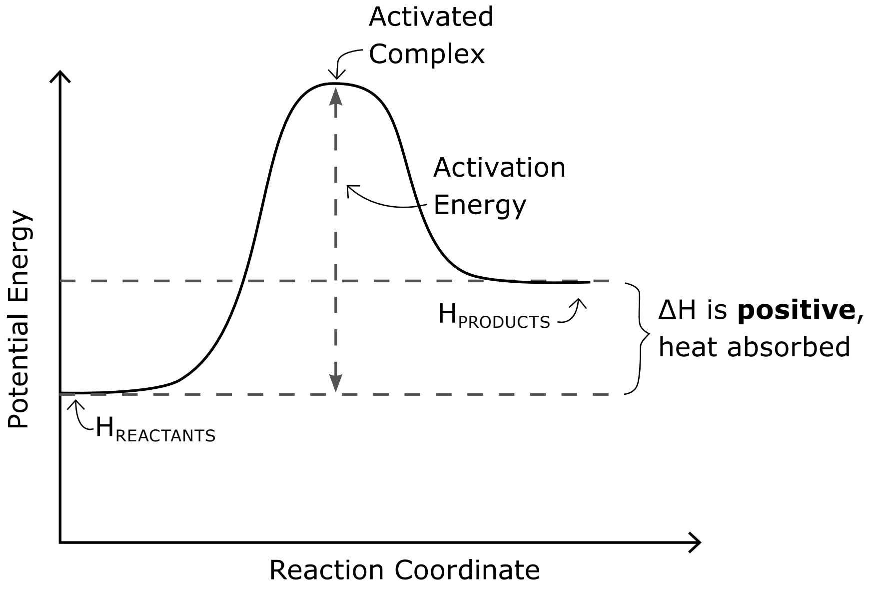
A reaction coordinate diagram represents the energy change that occurs during a chemical reaction. This type of diagram can also be called a **potential energy** (PE) or **reaction progress diagram**.

Exothermic Reaction



In the **exothermic reaction** shown above, the products possess less potential energy than the reactants. During the reaction, heat is lost from the system and ΔH is a negative value.

Endothermic Reaction



In an **endothermic reaction** (shown above), the products possess more potential energy than the reactants. This energy is absorbed from the surroundings, increasing the system's energy content, giving a positive ΔH value.

Reaction coordinate diagrams, or potential energy diagrams, provide a picture of the energy changes that occur as a chemical reaction proceeds. The reaction coordinate axis can be time or any other means of measuring reaction progress. The energy changes during a chemical reaction are much the same as those occurring during a roller-coaster ride. The beginning of a roller-coaster ride is a long, slow, uphill ride. Energy is expended by motors in the effort to get the cars to the top of the first hump. Once the cars reach the top of the first hump, it's all downhill from there. The cars have enough energy for the entire ride, regardless of the number of ups, downs, twists, and turns.

As reactant particles approach one another, they slow down due to repulsion between their valence electrons, and kinetic energy begins to convert into potential energy. If the collision between approaching particles is energetic enough, then the level of the potential energy will be raised to where an **activated complex** will form. The activated complex is the arrangement of the reacting molecules with partly broken and partly formed bonds at the peak of the barrier, which only exists for 10^{-13} seconds. At this point in time, called the **transition state**, the activated complex could form products or decompose back into reactants. If the activation energy is high, very few collisions will have enough energy to overcome the activation energy barrier and the reaction rate will be slow.

In a chemical reaction, energy is required to overcome the activation energy and to form the activated complex. (Just as with a roller-coaster ride, where if the cars do not make it to the top of a hump, the ride does not occur and there are several disappointed (or relieved) riders.) Once the activated complex is formed, the reaction could be successful and go to completion to form product. If particles do not possess the necessary activation energy to produce the activated complex, the reaction does not occur.

The image available at www.dlt.ncssm.edu/core/Chapter15-Kinetics/Chapter15-Animations/Molecular_collision_Ea.html shows the change in potential energy as two molecules collide.

You can also access this image/animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm#kinetics
2. scrolling down the page until you find the animation titled "**Molecular_collision_Ea.html**" or "**Molecular_collision_Ea.exe**"

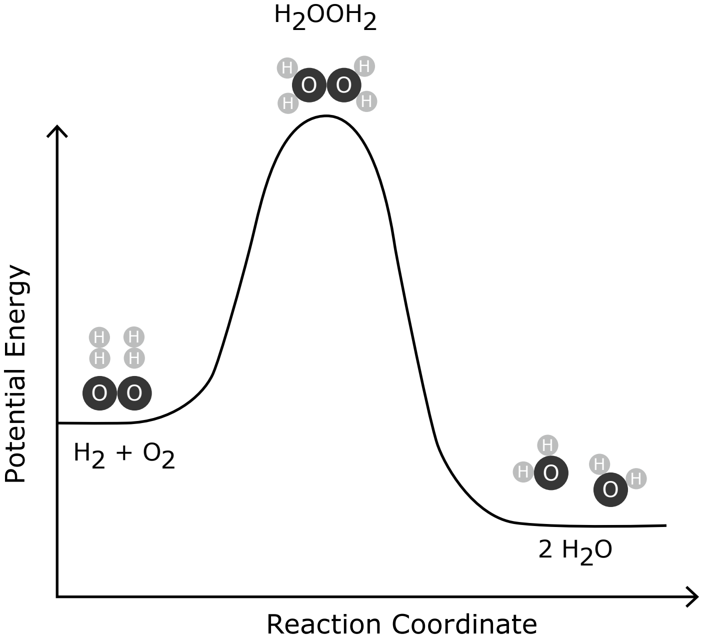
Notice that when particles of reactants have low kinetic energies, they cannot overcome the energy barrier to form products. The second part of the image illustrates the same reactant particles; however, in the second instance, they have higher kinetic energies and reach activation energy to form product.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Example 1

Let's look at the potential energy diagram for the formation of water from hydrogen and oxygen.

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

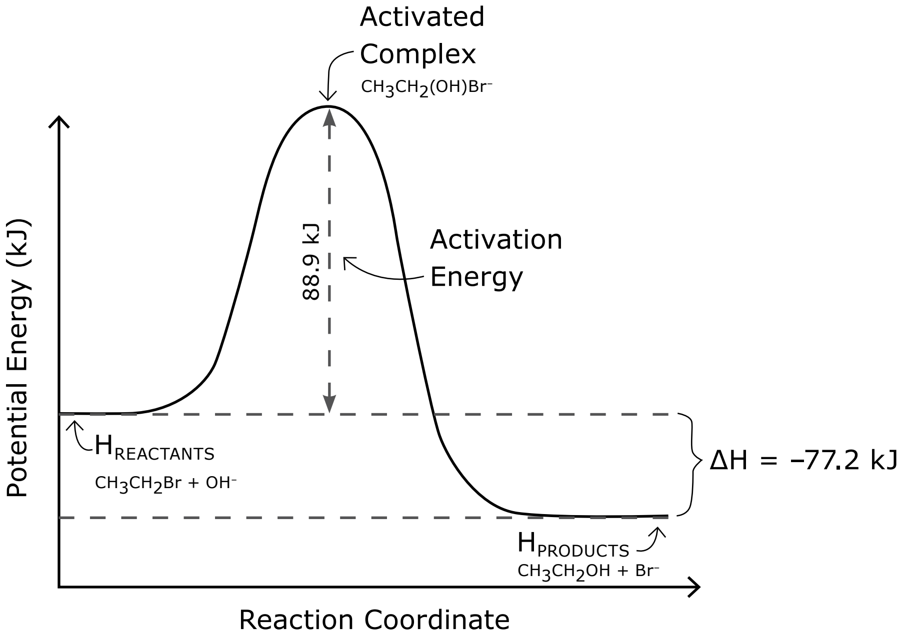


Hydrogen and oxygen can exist together without exploding or burning, due to the need to overcome activation energy. The activation energy is high because of the need to break covalent hydrogen-hydrogen bonds in hydrogen molecules, and oxygen-oxygen bonds in oxygen molecules. Notice that the transition state that occurs at the activated complex is a single particle. This H_2OOH_2 particle does not last very long, but must be formed for hydrogen to burn. You should also notice that the enthalpy of water is lower than the enthalpy of hydrogen and oxygen molecules. The energy lost in the formation of water is released in the form of heat and light energy. All combustion reactions are exothermic.

Example 2

Lets look at the reaction coordinate diagram for the following reaction:

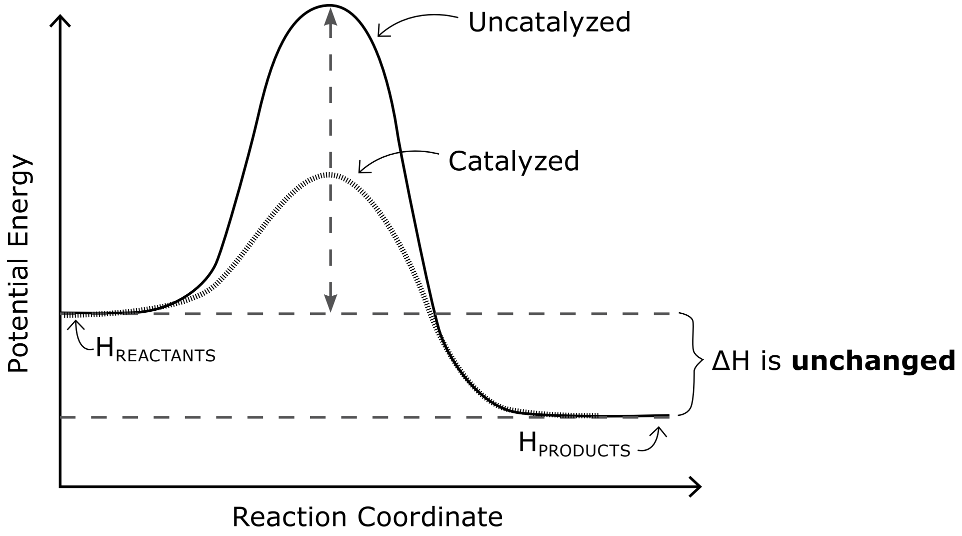
$$CH_3CH_2Br + OH^- \rightarrow CH_3CH_2OH + Br^-$$



- The activated complex in this reaction is $CH_3CH_2(OH)Br^-$. It is a single particle formed by all of the reactant particles.
- The activation energy is 88.9 kJ per mole of CH_3CH_2Br. Consequently, this reaction does not take place unless 88.9 kJ per mole of CH_3CH_2Br is added to the system.
- The enthalpy change is –77.2 kJ. This indicates an exothermic reaction.

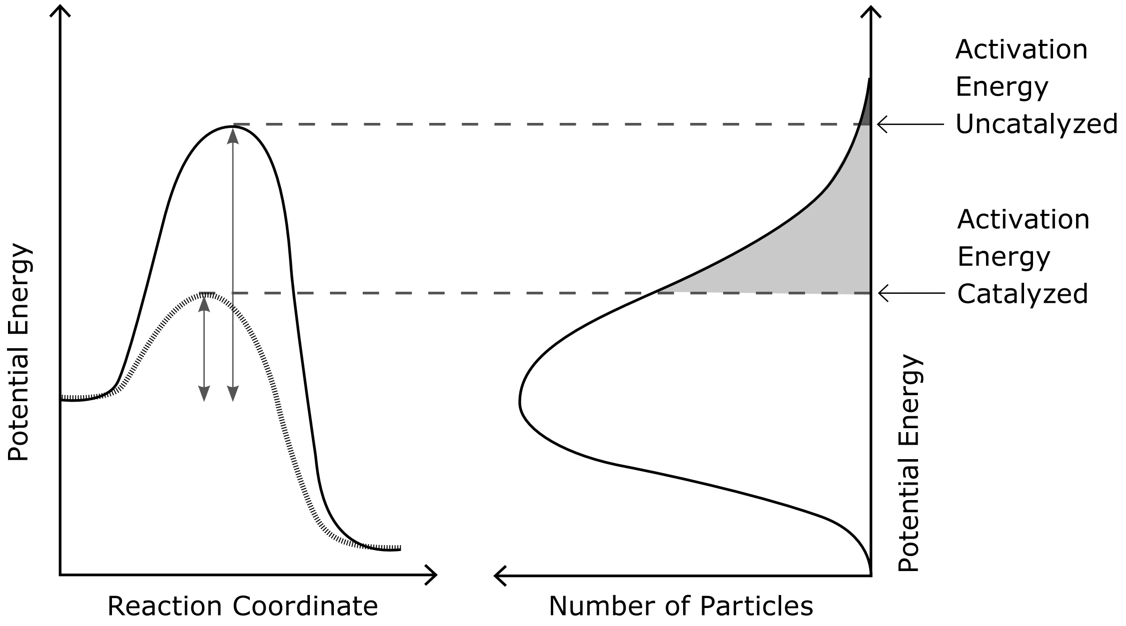
Effect of a Catalyst

A **catalyst** is a substance that speeds up or initiates a reaction without itself being permanently changed. Generally, a catalyst lowers the activation energy of a reaction. With a lower activation energy, previously ineffective collisions become more effective. As a result, the rate of the reaction increases. **Inhibitors** are the opposite of catalysts, stopping or slowing down reactions, by increasing the activation energy or interfering with collisions.



The above reaction coordinate diagram indicates that a catalyst speeds up a reaction by lowering the activation energy that the reaction must overcome. The catalyst does **not** affect the reaction products or the enthalpy change for the reaction: both remain the same. Catalysts will only provide an easier path for the reaction to proceed.

Lowering the required activation energy of a reaction increases the number of particles with enough energy to produce an effective collision. This means that more particles are likely to have the lower activation energy, resulting in effective collisions occurring more frequently. As a result, the reaction rate will increase.



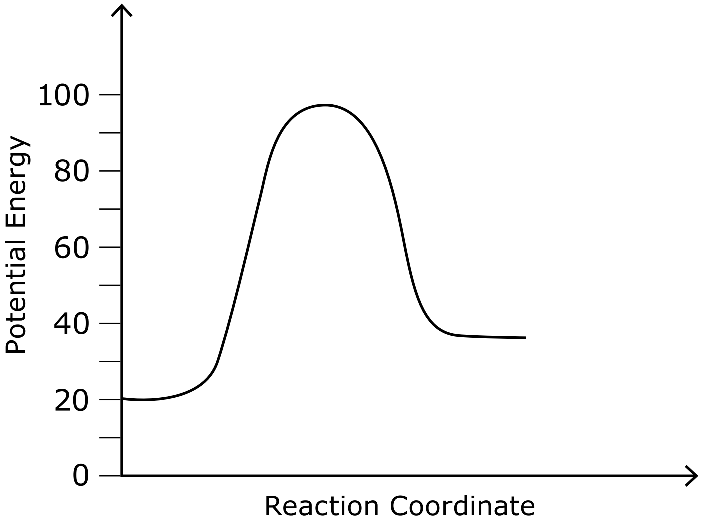
The shaded areas on the right side of the graph represent the number of particles with the required activation energy. Note that lowering the activation energy on the potential energy diagram corresponds to an increase in the number of particles with kinetic energies above the activation energy level. These particles can then be involved in a greater number of effective collisions, resulting in an increased reaction rate.

Reversible Reactions

As you will see in the unit on equilibrium, many reactions are reversible, meaning that they can also go in the opposite direction to what is written. The potential energy diagram for the reverse reaction is a mirror image to the forward reaction. In other words, a reaction with a positive ΔH will have a negative ΔH of the same value when reversed. A reaction with a negative ΔH will have, when reversed, a positive ΔH of the same value. The activation energy is the same for the forward and reverse reactions.

Learning Activity 3.6: Potential Energy Diagrams

1. Given the following reaction coordinate diagram:

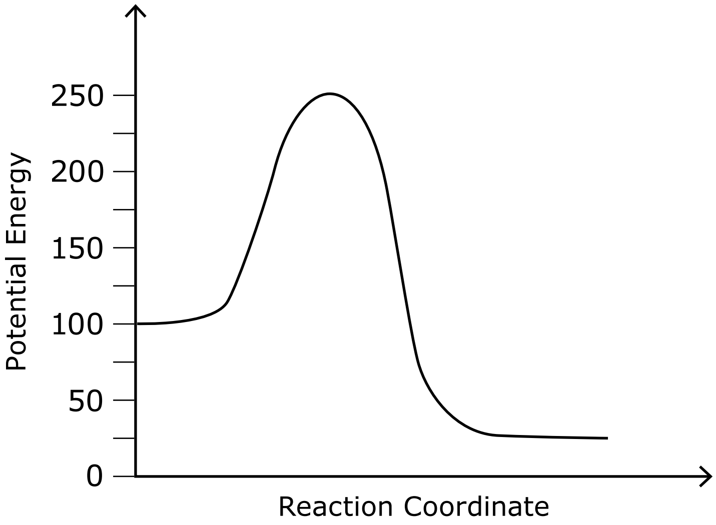


   a. What is the activation energy of the reaction shown by the diagram above?
   b. What is the enthalpy change for this reaction?
   c. Is this reaction endothermic or exothermic?

continued

Learning Activity 3.6: Potential Energy Diagrams (continued)

2. Given the following reaction coordinate diagram:



   a. What is the activation energy of the reaction in the diagram above?
   b. What is the enthalpy change for this reaction?
   c. Is this reaction endothermic or exothermic?
   d. What would be the activation energy of the *reverse* reaction?

3. What is the activated complex or transition state and how is it related to reaction rates?
4. Why does an exothermic reaction produce heat even though its energy value is negative?
5. Label each of these reactions as endothermic or exothermic.

   a. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat ______________________________
   b. $6C + 3H_2 \rightarrow C_6H_6$, $\Delta H = +83$ kJ ________________________________
   c. $2NO_2 + 112$ kJ $\rightarrow 2NO + O_2$ __________________________________
   d. $S + O_2 \rightarrow SO_2$, $\Delta H = -296$ kJ __________________________________

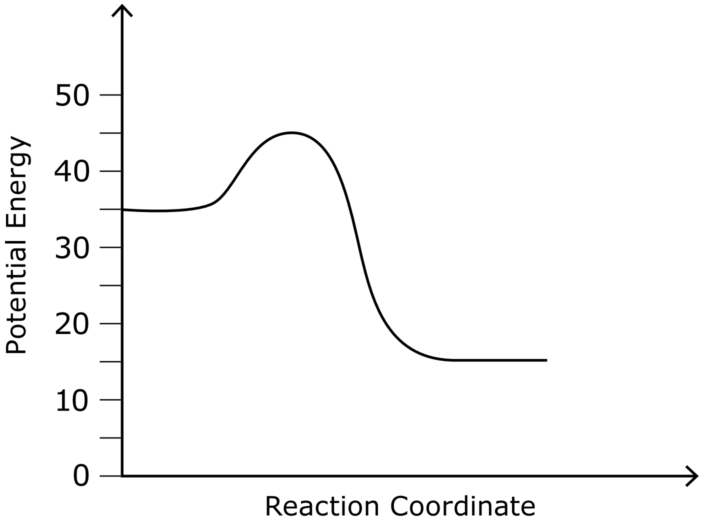
Check the answer key.

Lesson Summary

In this lesson, you learned how to draw potential energy diagrams for endothermic and exothermic reactions. You also learned how to identify the activated complex and the heat of reaction. In the next lesson, using the collision theory, you will identify factors which affect the rate of chemical reactions.

Assignment 3.3: Potential Energy Diagrams (10 marks)

1. Given the following reaction coordinate diagram:



a. What is the activation energy of the diagram above? *(2 marks)*

b. What is the enthalpy change for this reaction? *(2 marks)*

continued

Assignment 3.3: Potential Energy Diagrams (continued)

c. Is this reaction endothermic or exothermic? *(1 mark)*

d. What would be the activation energy of the reverse reaction? *(2 marks)*

e. Label the activated complex on the diagram. *(1 mark)*

2. Does every collision between reactant particles produce a reaction? Explain. *(2 marks)*

LESSON 7: REACTION MECHANISM (2 HOURS)

Lesson Focus

SLO C12-3-09: Explain the concept of a reaction mechanism.
Include: rate-determining step

Lesson Introduction

Very few chemical reactions occur as a result of a single, effective collision. Most reactions happen as a series of smaller reactions between two particles at a time, where each step is a simple event involving simple collisions. The sum of all of these simpler steps results in the overall reaction. This series of steps is called a **reaction mechanism**.

Which Collisions Result in Reactions?

Remember that a collision between molecules can have one of two results:

- If the colliding molecules do not have enough energy to surpass the activation energy, and the molecules simply rebound and move away from each other, no chemical reaction occurs.
- If the colliding molecules have enough energy to surpass the activation energy, the molecules combine to form an activated complex at the **transition state**.

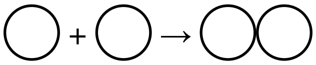
Molecules in this transition state are very unstable and are just as likely to return to reactants as they are to go to products. As you learned in a previous lesson, the collision must have enough energy to surpass the activation energy because this is the amount of energy required to break the bonds in reactants and form new ones in products. If the collision is not energetic enough to surpass the activation energy barrier, no products will form. If the collision is energetic enough to surpass the activation energy barrier, an activated complex will form—half of all activated complexes form products, while the other half return to reactants.

Reaction Mechanisms

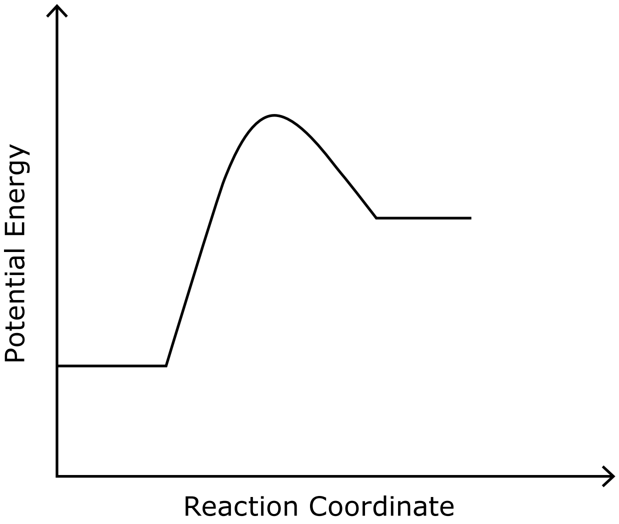
In an **elementary reaction**, the reactants are converted to products in a single step. Only two particles will collide at one time.

In an elementary reaction such as

$$H + H \rightarrow H_2$$



there is only one activated complex and the potential energy curve appears as follows:



However, most reactions consist of a number of elementary reactions. For the overall reaction 2 NO + O_2 → 2 NO_2, there is an intermediate step.

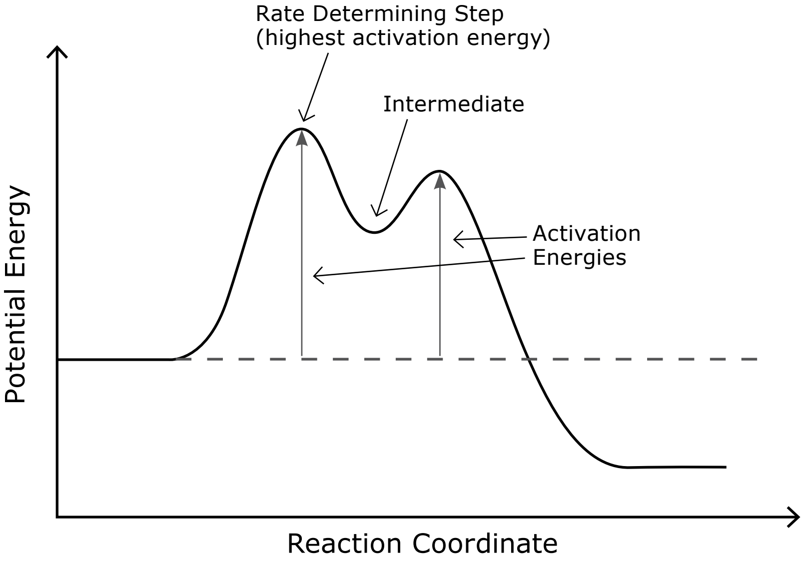
$NO + NO \rightarrow N_2O_2$ Step 1

$N_2O_2 + O_2 \rightarrow 2\,NO_2$ Step 2

$2NO + O_2 \rightarrow 2\,NO_2$ Overall

(N_2O_2 produced in Step 1 is consumed in Step 2)

An **intermediate** is the product of a reaction that immediately becomes a reactant in another reaction. A **reaction mechanism** lists the series of elementary reactions that take place during the course of a complex reaction. *Each step of the reaction has an energy requirement.* If there are multiple steps, the potential energy diagram will have several hills and valleys (one for each intermediate step).



The hills correspond to the energy levels of the activated complexes.

The intermediate products are found in the valleys of the potential energy diagrams. These products will be used as reactants in the next step of the reaction. Therefore, intermediates do not appear in the final chemical equation.

One of the steps in the reaction will always be slower than the others: it is called the **rate-determining step**. The rate determining step can either be indicated as RDS or referred to as the slow step. On a potential energy diagram, it is always the step in the reaction with the highest activation energy.

Catalysts, such as enzymes or platinum, first appear on the left-hand side of the mechanism and then appear on the right-hand side in a lower step of the mechanism. Remember that catalysts are recycled and can be used more than once. A catalyst is cancelled out and does not appear in the net reaction.

Example 1

For the following reaction mechanism:

1. Identify the intermediates.
2. Identify the rate determining step.
3. Write the overall equation.
4. Indicate whether the overall reaction is endothermic or exothermic.

$A + B \rightarrow I_1 + G$

$I_1 + D \rightarrow I_2 + F$ (slow)

$I_2 + H \rightarrow G + K$, $\Delta H = +15$ kJ

Solution:

1. The intermediates are used as reactants in the next step of the reaction: I_1 and I_2.
2. The rate determining step is the slow step: $I_1 + D \rightarrow I_2 + F$.
3. The overall equation is $A + B + D + H \rightarrow 2G + F + K$.
4. $\Delta H = +15$ kJ; therefore, the reaction is endothermic.

Example 2

Given the following mechanism:

Step 1 (RDS): $HBr + O_2 \rightarrow HOOBr$ (slow)

Step 2: $HOOBr + HBr \rightarrow 2HOBr$ (fast)

Step 3: $2HOBr + 2HBr \rightarrow 2H_2O + 2Br_2$ (fast)

Net reaction: $4HBr + O_2 \rightarrow 2H_2O + 2Br_2$

Solution:

1. The reactants are 4HBr and O_2.
2. The products are $2H_2O$ and $2Br_2$.
3. The intermediates do not appear in the final chemical equation: HOOBr and 2HOBr.

Learning Activity 3.7: Describing Reaction Mechanism

Answer each of the following in the space provided.

1. Consider the following reaction:
   $X_{(s)} + Y_{(g)} + \text{heat} \rightarrow Z_{(aq)}$

   Suggest ways to increase the reaction rate of this chemical reaction. List all possibilities.

2. For each of the following reaction mechanisms:
   - Identify the intermediates.
   - Identify the rate determining step.
   - Determine the overall equation.

   a. $H_2 + ICl \rightarrow HI + HCl$ (slow)
   $HI + ICl \rightarrow I_2 + HCl$

   b. $O_3 \rightarrow O_2 + O$ (slow)
   $O_3 + O \rightarrow 2\,O_2$

3. Describe where each of these would appear in a reaction mechanism:
   - Reactants
   - Products
   - Intermediates
   - Catalysts

4. Describe a reaction mechanism.
5. How are the elementary steps in a reaction related to the overall reaction mechanism?

Check the answer key.

Lesson Summary

In this lesson, you learned that most chemical reactions do not occur in one single step. An intermediate is the product of one reaction that immediately becomes a reactant in another reaction. A reaction mechanism lists the series of elementary reactions that take place during the course of a complex reaction. The rate determining step will always be slower than the others.

Assignment 3.4: Reaction Mechanisms (17 marks)

1. For each of the following reaction mechanisms: *(10 marks)*
   - identify the intermediates *(2 marks each)*
   - identify the rate determining step *(1 mark each)*
   - determine the overall equation *(2 marks each)*

   a. $2\,NO \rightarrow N_2O_2$
   $N_2O_2 + H_2 \rightarrow N_2O + H_2O$ (slow)
   $N_2O + H_2 \rightarrow N_2 + H_2O$

   b. $I^- + S_2O_8^{2-} \rightarrow IS_2O_8^{3-}$ (slow)
   $IS_2O_8^{3-} \rightarrow 2\,SO_4^{2-} + I^+$
   $I^+ + I^- \rightarrow I_2$
   $I_2 + I^- \rightarrow I_3$

continued

Assignment 3.4: Reaction Mechanisms (continued)

2. Given the following mechanism:

   Step 1 (RDS): ? → ?

   Step 2: $Cl + H_2 \rightarrow HCl + H$

   Step 3: $H + Cl_2 \rightarrow HCl + Cl$

   Step 4 $Cl + Cl \rightarrow Cl_2$

   Overall: $H_2 + Cl_2 \rightarrow 2HCl$

   Write the reaction for Step 1 and identify the/all intermediate(s) and the catalyst. *(5 marks)*

continued

Assignment 3.4: Reaction Mechanisms (continued)

3. Differentiate between a balanced chemical equation and a reaction mechanism. *(2 marks)*

Notes

LESSON 8: RATE LAW AND ORDER (3 HOURS)

Lesson Focus

SLO C12-3-10: Determine the rate law and order of a chemical reaction from experimental data.
Include: zero-, first-, and second-order reactions and rate-versus-concentration graphs

Lesson Introduction

The rate of a reaction for gases, liquids, and solutions depends on several factors, one of which is the concentration of reactants. This relationship can be expressed mathematically by the rate law, which you will learn about in this lesson. You will also determine which reactants have a greater effect on reaction rate by observing the relationship between each concentration and the rate.

Calculating Average Rate

The rate of a reaction is often described in terms of the change in concentration of the reactants or the products. The **average rate** of a reaction is the change in concentration (or other property) over a period of time. For the equation

$$\mathrm{A} \rightarrow \mathrm{B}$$

the average rate can be expressed mathematically in both of the following ways:

$$\text{average rate} = \frac{\Delta[\mathrm{B}]}{\Delta\text{time}}$$

$$\text{average rate} = \frac{\Delta[\mathrm{A}]}{\Delta\,\text{time}}$$

Chemists often use the square brackets to mean "the concentration of" the substance in the brackets. The Δ, or the Greek letter delta, means "change in." You can write this equation in another way, as well.

$$\text{average rate} = \frac{\Delta[\text{A}]}{\Delta\text{time}}$$

$$= \frac{(\text{final concentration of A}) - (\text{initial concentration of A})}{\text{final time} - \text{initial time}}$$

$$= \frac{[\text{A}]_{\text{final}} - [\text{A}]_{\text{initial}}}{t_{\text{final}} - t_{\text{initial}}}$$

Example 1:

According to the reaction A → B, the following data was collected:

Time (s)	Concentration of B (mol/L)
0.0	0.00
10.	0.30
20.	0.50
30.	0.60
40.	0.65
50.	0.67

1. What is the average rate over the entire 50. seconds? Remember that rate is a change over time.

$$\text{average rate} = \frac{\Delta[\text{B}]}{\Delta t} = \frac{[\text{B}]_{50\text{s}} - [\text{B}]_{0\text{s}}}{50.\text{s} - 0.0\,\text{s}}$$

$$= \frac{0.67\text{ mol/L} - 0.0\text{ mol/L}}{50.\text{s} - 0.0\,\text{s}} = 0.013\text{ mol/L} \cdot \text{s}$$

Note that the unit for rate is mol/L·s. You should think of this as "moles per litre every second."

2. What is the average rate for the interval 20. s to 40. s? The units for rate can be written as mol/L•s or mol $L^{-1}s^{-1}$.

$$\text{average rate} = \frac{\Delta[B]}{\Delta t} = \frac{[B]_{40s} - [B]_{20s}}{40.s - 20.s}$$

$$= \frac{0.65 \text{ mol/L} - 0.50 \text{ mol/L}}{40.s - 20.s} = 0.0075 \text{ mol/L} \cdot \text{s}$$

Example 2:

The decomposition of nitrogen dioxide produces nitrogen monoxide and oxygen according to the following reaction:

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

The following data has been collected:

Time (s)	$[NO_2]$ (mol/L)	[NO] (mol/L)	$[O_2]$ (mol/L)
0.0	0.100	0.00	0.00
100.	0.066	0.034	0.017
200.	0.048	0.052	0.026
300.	0.038	0.062	0.031
400.	0.030	0.070	0.035

Calculate the average rate of decomposition of NO_2 over 400 s.

$$\text{average rate} = \frac{\Delta[NO_2]}{\Delta t} = \frac{[NO_2]_{400s} - [NO_2]_{0s}}{400.s - 0.0s}$$

$$\frac{0.030\,\text{mol/L} - 0.100\,\text{mol/L}}{400.s - 0.0s} = 1.8 \times 10^{-4} \text{mol/L} \cdot \text{s}$$

Note that the calculated rate is a negative number because reactants are used up. Rate is always expressed as a *positive* value. The actual value of the average rate is 1.8×10^{-4} mol/L·s.

Rewrite the equation for the average rate of a reactant with a negative sign:

$$\text{rate} = -\frac{\Delta[NO_2]}{\Delta t}$$

Rate Constant and the Rate Law

You learned in the previous section that the rate at which reactant forms product can be expressed as the change in the concentration of a reactant over time. Therefore, the rate of disappearance of a reactant must be, under constant temperature and pressure, proportional to its concentration:

$$\text{Rate} \propto [A]$$

Since proportionality statements do not indicate the degree of the relationship, we insert a constant, k, that will allow us to make mathematical calculations. In this case, we call ***k*** the **rate constant**.

$$\text{Rate} = k\ [A]$$

The rate constant is a specific value that is experimentally determined for each reaction and applies only to a given temperature and pressure. With every change in conditions (other than the concentration of reactants), a new value for the rate constant must be determined. A large value ***k*** indicates that the reaction proceeds quickly and product forms rapidly. If the value for ***k*** is low, then product forms slowly.

In reactions that have more than one reactant, the rate of reaction likely will be affected by the concentration of all of the reactants. The general equation used to express the rate of the reaction is called the **rate law** and is written as follows:

General reaction: $aA + bB \rightarrow cC + dD$

Rate Law: $\text{Rate} = k\ [A]^{m}\ [B]^{n}$

Where ***m*** and ***n*** are exponents that we call the **reaction order**, the values of ***m*** and ***n*** can only be determined experimentally.

Reaction Order

To this point, you have learned that changing the concentration of substances taking part in a reaction usually changes the rate of the reaction, and that the rate law is a mathematical statement of that relationship. It makes sense that there is a direct relationship between the concentration of reactants and the rate of a reaction—that doubling the concentration of a reactant should mean the reaction will go twice as fast (though it sometimes does not). This is where the rate law becomes important in understanding how a reaction will proceed—and the values for ***m*** and ***n*** are the key to that understanding.

First-Order Reactions

Let's start with a simple reaction where there is a single reactant, such as the decomposition of hydrogen peroxide.

$$H_2O_2 \rightarrow H_2 + O_2$$

Experimentally, we find that changing the concentration of the hydrogen peroxide has a direct, linear relationship with the rate of reaction. Doubling the H_2O_2 concentration doubles the reaction rate. Halving the concentration halves the rate. Increasing the concentration by 75% increases the rate by 75%. We call this 1-to-1 relationship a first-order reaction. The rate law looks like this:

$$\text{Rate} = k\ [H_2O_2]^1$$

Note that the exponent is 1. Remember that any value to the power of 1 is the same value. Mathematically, any change in H_2O_2 concentration creates the same degree of change in rate.

Second-Order Reactions

A different situation occurs in the decomposition of nitrogen dioxide.

$$2NO_2 \rightarrow 2NO + O_2$$

In this case, doubling the concentration of NO_2 causes a quadrupling of the rate of production of nitric oxide and oxygen. Tripling the NO_2 concentration causes the rate to increase 9 times. This is an exponential relationship where a change in the reactant concentration causes a squared change in rate. The rate law is written as

$$\text{Rate} = k\,[NO_2]^2$$

This is called a second-order reaction.

Zero-Order Reactions

There are some reactions in which a change in the reactant concentration has no effect on the rate. The reverse of the famous Haber process is one example.

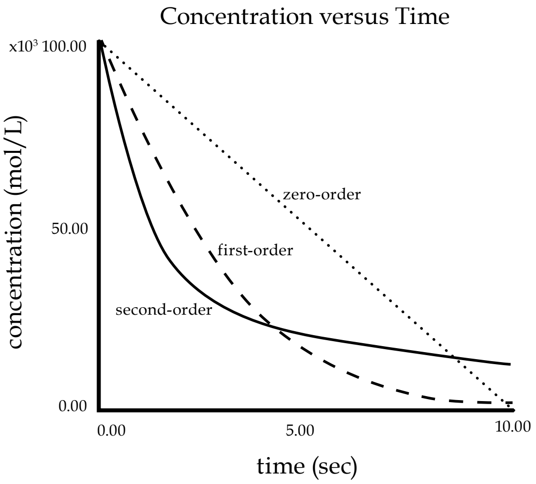
$$2NH_{3(g)} \rightarrow 3H_{2(g)} + N_{2(g)}$$

In this reaction, there is no change in the rate of the reaction regardless of the concentration of the ammonia gas. The rate law is written as

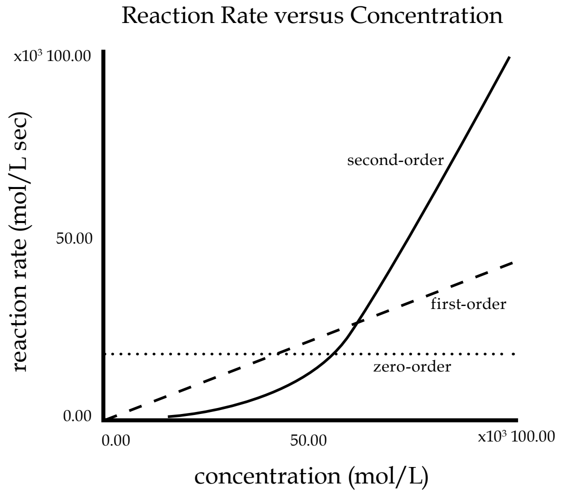
$$\text{Rate} = k\,[NH_3]^0$$

Mathematically, any value raised to a power of 0 is equal to 1. Although we could simply write *rate* = *k*, we usually write in the reactant in order to understand that there is a reaction occurring. As we shall see, if there is more than one reactant, then we can leave off any reactants that are zero-order from the rate law expression.

These three scenarios are illustrated in the following graphs. The graph at the left shows concentration-time plots for zero-order, first-order, and second-order reactions. The corresponding rate-concentration plots are shown at the right.



Concentration versus Time



Reaction Rate versus Concentration

In examining the graphs, notice that, as the reaction progresses, the concentration of reactant decreases. Examine the graphs and take note of the following points:

- For a **zero-order reaction**, the rate of reaction is constant as the reaction progresses.
- For a **first-order reaction**, the rate of reaction is directly proportional to the concentration. As the reactant is consumed during the reaction, the concentration drops and so does the rate of reaction.
- For a **second-order reaction**, the rate of reaction increases with the square of the concentration, producing an upward curving line in the rate-concentration plot. For this type of reaction, the rate of reaction decreases rapidly as the concentration of the reactant decreases.

Orders of reaction are always found by doing an experiment. This means that you can't determine the order of a reaction just by looking at the balanced chemical equation for the reaction. Suppose that you had done some experiments to find out what happens to the rate of a reaction as the concentration of one of the reactants changes for the following:

$$A + B \rightarrow C + D$$

While both reactants affect reaction rate, let's concentrate on A for this example. Some of the simple things that you might find are as follows:

1. **The rate of reaction is proportional to the concentration of A.** This means that if you double the concentration of A, the rate doubles as well. If you increase the concentration of A by a factor of four, the rate also increases by four times. Earlier in this lesson, you learned to express this relationship using symbols:

   $$\text{Rate} \ \alpha \ [A]^1$$

   You can also write this as the rate law by inserting the constant, *k*.

   $$\text{Rate} = k\,[A]^1$$

2. **The rate of reaction is proportional to the square of the concentration of A.** This means that if you doubled the concentration of A, the rate would go up four times. If you tripled the concentration of A, the rate would increase nine times. In symbolic terms,

   $$\text{Rate} \ \alpha \ [A]^2$$

   $$\text{Rate} = k\,[A]^2$$

By doing experiments involving a reaction between A and B, you would find that the rate of the reaction was related to the concentrations of A and B in the following way:

$$\text{Rate} = k\,[A]^m\,[B]^n$$

The concentrations of A and B have to be raised to some power to show how they affect the rate of the reaction. These powers are called the orders of reaction, with respect to A and B. Remember that if the order of reaction with respect to A is 0 (zero), it would not be included in the rate equation because any value raised to the power of 0 is equal to 1. As you see in many different scenarios, if the exponent is 1, you do not need to write it; however, doing so helps you remember to include it when you are asked to calculate overall orders.

Order with Respect to a Reactant

As previously stated, reaction rates depend in part on the concentration of the reactants. Consider the following reaction:

$$H_{2(g)} + I_{2(g)} \rightarrow 2\,HI_{(g)}$$

If the concentration of iodine is doubled and the concentration of hydrogen is kept the same, the number of collisions between hydrogen and iodine molecules doubles. Consequently, the reaction rate also doubles.

$$\text{Reaction rate } \alpha \ [I_2]$$

The same results occur if the concentration of hydrogen is doubled.

$$\text{Reaction rate } \alpha \ [H_2]$$

This means the reaction rate depends on *both* the concentrations of hydrogen and iodine. If both the hydrogen and iodine concentrations are doubled, there are four times as many molecules and, therefore, four times as many collisions. The reaction rate consequently quadruples.

$$\text{Reaction rate } \alpha \ [I_2]^1 \ [H_2]^1$$

The reaction rate depends on the product of the concentrations of hydrogen and iodine. This information is summarized in the rate law expression

$$\text{Rate} = k\,[I_2]^1\,[H_2]^1$$

In this reaction, we would state the following:

The order with respect to I_2 is first order, and the order with respect to H_2 is also first order.

Overall Order

The overall order of the reaction is found by adding up the individual orders. For example, if the reaction is first order with respect to both A and B ($m = 1$ and $n = 1$), the overall order is 2. This an overall second-order reaction. For the following examples, continue to use $A + B \rightarrow C + D$ as the chemical reaction.

Example 1

$$\text{Rate} = k\,[A]^0\,[B]^2 \text{ which becomes simply:}$$

$$\text{Rate} = k\,[B]^2$$

The order with respect to A is 0. The order with respect to B is 2: it's a second-order reaction with respect to B. The reaction is also second order overall (because $0 + 2 = 2$).

Example 2

Rate = $k\,[A]^1\,[B]^0$, which becomes simply:

Rate = $k\,[A]$ or Rate = $k\,[A]^1$

This reaction is first order with respect to A, and zero order with respect to B because the concentration of B does not affect the rate of the reaction. The reaction is, therefore, first order overall (because $1 + 0 = 1$).

What if you have more than two reactants? The number of reactants is irrelevant. The concentration of each reactant will be included in the rate equation, raised to some power. Those powers are the individual orders of reaction. The overall order of the reaction is still found by adding them all up.

Example 3

A reacts with B according to the following balanced chemical equation:

$$A + 3B \rightarrow 2C$$

Rate = $k\,[A]^2\,[B]^3$

The order of the reaction in each reactant is the value of the exponent associated with the reactant. As such, the following applies to the previous reaction:

- The order with respect to A is 2, and the order with respect to B is 3.
- The overall order of the reaction is the sum of the exponents for each reactant (5th).

Rate Law Derived from Experimental Data

Accurate rate laws can only be determined from experimental data. In this section, you will be presented with the data from an experiment in which several trials of different concentrations of reactants were tested. The initial rates of reaction for each trial will have been determined. By analyzing the data, we can determine the rate law.

In order to determine the effect of a single reactant concentration on the reaction rate, you must choose a set of data where the other reactant remains constant. Then, evaluate the rate of concentration change and the effect it has on the reaction rate.

Example 1

Consider the following reaction.

$$NH_4^+{}_{(aq)} + NO_2^-{}_{(aq)} \rightarrow N_{2(g)} + 2H_2O_{(l)}$$

Using the data in the table below, answer the questions that follow.

Trial	Initial $[NO_2^-]$(mol/L)	Initial $[NH_4^+]$(mol/L)	Initial Rate (mol/L·s)
1	0.0100	0.200	5.4×10^{-7}
2	0.0200	0.200	10.8×10^{-7}
3	0.0400	0.200	21.5×10^{-7}
4	0.0600	0.200	32.3×10^{-7}
5	0.200	0.0202	10.8×10^{-7}
6	0.200	0.0404	21.6×10^{-7}
7	0.200	0.0606	32.4×10^{-7}
8	0.200	0.0808	43.3×10^{-7}

1. Determine the order of the reaction with respect to each of the reactants.

   Order with respect to NO_2^-:

   We can compare Trial 1 and Trial 2 data because, during this time, the concentration of the ammonium does not change, but that of the nitrite ion does. And so, any change noted can only be due to the change in NO_2^- concentration.

   What do you observe? Doubling the concentration of NO_2^- also doubles the reaction rate. Since $2^1 = 2$, we can assign the order with respect to NO_2^- to be 1.

   Order with respect to NH_4^+:

   We can compare Trial 5 and Trial 6 data because, during this time, the concentration of NO_2^- remains constant. As such, any change noted can only be due to the change in concentration of NH_4^+.

   What do you observe? Doubling the concentration of NH_4^+ also doubles the reaction rate. Since $2^1 = 2$, we can assign the order with respect to NH_4^+ to be 1.

2. Write the rate law for this reaction.

   Using this information, we can assess the rate law to be

   $$Rate = k[NO_2^-]^1\,[NH_4^+]^1$$

3. Determine the overall order of the reaction.

   The overall order would be 1 + 1 = 2nd order.

4. We can even calculate k for this reaction at this temperature.

   Pick one trial and insert the values into the rate law you just determined.

   Using data from Trial 1:

   $Rate = k[NO_2^-]^1\ [NH_4^+]^1$

   Rearranging:

   $$k = \frac{Rate}{[NO_2^-]^1\ [NH_4^+]^1} = \frac{5.4 \times 10^{-7}\ mol/L \cdot s}{(0.0100 mol/L)^1 (0.200 mol/L)^1} = 2.7 \times 10^{-4}$$

   The rate constant for this reaction at this temperature is 2.7 x 10-4.

Example 2

For the reaction $NO_{2(g)} + CO_{(g)} \rightarrow NO_{(g)} + CO_{2(g)}$, the following data was obtained. Determine the overall rate law for this reaction.

Trial	Initial Rate (mol/L · s)	Initial [NO_2] (mol/L)	Initial [CO] (mol/L)
1	0.0050	0.10	0.10
2	0.080	0.40	0.10
3	0.0050	0.10	0.20

Solution:

1. Take the ratio of the initial rates for the two trials in which only one reactant is changed.

   $$\frac{\text{Trial 2 } [NO_2]}{\text{Trial 1 } [NO_2]} = \frac{0.40}{0.10} = 4 \text{ times (quadrupled the concentration)}$$

   $$\frac{\text{Trial 2 rate}}{\text{Trial 1 rate}} = \frac{0.080}{0.0050} = 16 \text{ times (rate increased 16 times)}$$

   By increasing the concentration four times, the reaction time is increased by 16. Since $4^2 = 16$, this means that the rate depends on the square of the concentration of NO_2. The reaction is second order with respect to NO_2. The rate law would be rate $= k[NO_2]^2$.

2. Take the ratio of the initial rates for Trials 1 and 3. These trials have the concentration of CO changing.

$$\frac{\text{Trial 3 [CO]}}{\text{Trial 1 [CO]}} = \frac{0.20}{0.10} = 2 \text{ times (doubled the concentration)}$$

$$\frac{\text{Trial 3 rate}}{\text{Trial 1 rate}} = \frac{0.0050}{0.0050} = 1 \text{ time (rate does not increase)}$$

By increasing the concentration of CO, the experimental data shows that the reaction rate does not change. The concentration of CO does not change the rate of reaction; therefore, because $2^0 = 1$, the reaction is zero order with respect to CO.

The rate law would be rate $= k[NO_2]^2[CO]^0 = k[NO_2]^2(1) = k[NO_2]^2$.

Learning Activity 3.8: Understanding Rate Law

1. For the reaction A → products, the following data was collected:

Time (min)	Mass of A (g)
0.0	25.0
1.0	20.0
2.0	17.0
3.0	15.0
4.0	13.0
5.0	12.0

   a. Calculate the average rate, in g A/min, over the entire 5 minutes.
   b. Calculate the average rate for the interval between 2.0 and 4.0 minutes.
   c. Calculate the average rate for the interval between 3.0 and 4.0 minutes.

2. The decomposition of acetaldehyde to methane and carbon dioxide occurs according to the following equation:

$$CH_3CHO_{(g)} \rightarrow CH_{4(g)} + CO_{(g)}$$

continued

Learning Activity 3.8: Understanding Rate Law (continued)

The following are the results of an experiment:

Time (s)	$[CH_3CHO]$ (mol/L)
42	0.00667
73	0.00626
105	0.00586
190	0.00505
242	0.00464
310	0.00423
384	0.00383
480	0.00342
665	0.00282
840	0.00241

2. a. Calculate the rate of decomposition of acetaldehyde between 42 s and 105 s.
   b. Calculate the rate of decomposition in the interval 190 s to 480 s.
3. A first-order reaction initially proceeds at a rate of 0.5 mol/L·s. What is the rate when *half* the starting material remains? When *a quarter* of the starting material remains?
4. How does the magnitude of the k value relate to the speed of the reaction?
5. The rate law for the following reaction is second order and first order with respect to NO. Write the rate law expression for this reaction.

   $$NO_{(g)} + O_{3(g)} \rightarrow NO_{2(g)} + O_{2(g)}$$

6. The rate law for the reaction of nitrogen and fluorine gases is as follows:

   $$\text{Rate} = k\ [NO_2]^1 [F_2]^1$$

   What will happen to the reaction rate if the F_2 concentration is tripled? Show your work.

continued

Learning Activity 3.8: Understanding Rate Law (continued)

7. Use the equation and data table below to answer the following questions.

$$NO_{(g)} + O_{2(g)} \rightarrow NO_{2(g)}$$

Experiment	**Initial $[O_2]$ (mol/L)**	**Initial [NO] (mol/L)**	**Initial Reaction Rate (mol/dm³·s)**
1	0.0010	0.0010	7.10
2	0.0040	0.0010	28.4
3	0.0040	0.0030	256
4	0.0020	0.0010	14.1

a. Determine the rate law for the reaction.

b. Determine the overall order of the reaction.

Check the answer key.

Lesson Summary

In this lesson, you learned that the rate of a reaction can be described in terms of the change in concentration of the reactants or the products over time. A rate law equation displays this relationship mathematically. The rate law expression also shows the orders of reaction, which are always found experimentally.

Assignment 3.5: Rate and Order (15 marks)

1. Below is the data from an experiment that studied the following reaction:

$$2\ HCl_{(aq)} + CaCO_{3(s)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

HCl was placed in a beaker and massed immediately after adding $CaCO_3$ chips (time = 0). The mass of the beaker was recorded at 1.0 minute intervals for a total of 10 min. The loss of mass is the amount of carbon dioxide gas that escapes from the beaker.

Time (min)	**Mass of Beaker and Contents (g)**	**Mass Loss (CO_2 Produced) (g)**
0.0	200.00	0.00
1.0	199.40	0.60
2.0	199.00	1.00
3.0	198.65	1.35
4.0	198.35	1.65
5.0	198.10	1.90
6.0	197.90	2.10
7.0	197.75	2.25
8.0	197.65	2.35
9.0	197.57	2.43
10.0	197.52	2.48

a. Determine the average rate, in g of CO_2/min, over the entire 10 minutes. *(2 marks)*

b. Determine the average rate for the following intervals.

First 3.0 minutes: *(2 marks)*

6.0 to 10.0 minutes: *(2 marks)*

continued

Assignment 3.5: Rate and Order (continued)

2. Assume that SO_2 and O_2 react according to the rate law below.

   $Rate = k[SO_2]^2[O_2]^1$

   How does the rate change if:

   a. The concentration of SO_2 is doubled? Show your rationale. *(2 marks)*

   b. The concentration of O_2 is tripled? Show your rationale. *(2 marks)*

3. For the reaction $H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$, the following data was obtained.

Experiment	Initial $[H_2]$	Initial $[I_2]$	Initial Rate of Formation of HI
1	2.0 mol/L	2.0 mol/L	0.20 mol/dm^3/s
2	2.0 mol/L	4.0 mol/L	0.40 mol/dm^3/s
3	4.0 mol/L	4.0 mol/L	0.80 mol/dm^3/s

   a. Write the rate law for this reaction. Show your work. *(4 marks)*

   b. Determine the overall order of the reaction. *(1 mark)*

MODULE 3 SUMMARY

In Module 3, you learned that chemistry focuses largely on chemical reactions. Some reactions, such as rusting, occur very slowly, while others, such as explosions, occur very quickly. **Kinetics** is the branch of chemistry that studies the speed or **rate** at which chemical reactions occur. Many chemical reactions occur in more than one step. In this study of kinetics, you learned about the steps, or the **mechanism**, of a chemical reaction. It is important for chemists to understand the mechanism of a reaction and the factors that affect its rate. With this knowledge, chemists can control a chemical reaction in order to maximize the manufacturing of commercial products.

You have now completed Module 3. You have almost completed half of the course! It is now time to submit your assignments and then write your midterm examination. Read the following directions and double-check that you have submitted all the required work. Then, read the details regarding the midterm examination.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 3 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:

☐ Cover Sheet for Module 3 (found at the end of the course Introduction)
☐ Assignment 3.1: Reaction Rate
☐ Assignment 3.2: Collision Theory and Reaction Rate
☐ Assignment 3.3: Potential Energy Diagrams
☐ Assignment 3.4: Reaction Mechanisms
☐ Assignment 3.5: Rate and Order

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: pens, some blank paper, a ruler, and a scientific calculator.

A maximum of **3 hours** is available to complete your midterm examination. When you have completed it, the proctor will then forward it to the Distance Learning Unit for assessment. Good luck!

Midterm Practice Examination and Answer Key

To help you succeed in your examination, a practice examination can be found in the learning management system (LMS). The midterm practice examination is very similar to the actual examination that you will be writing. The answer key is also included so that, when you have finished writing the practice examination, you can check your answers. This will give you the confidence that you need to do well on your examination. If you do not have access to the Internet, contact the Distance Learning Unit at 1-800-465-9915 to get a copy of the practice examination and the answer key.

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GRADE 12 CHEMISTRY (40S)

Module 3
Chemical Kinetics

Learning Activity Answer Keys

MODULE 3:
CHEMICAL KINETICS

Learning Activity 3.1: Measuring Reaction Rate

1. For each of the following ways of determining reaction rate, state the units of measurement.

   a. Temperature change over time: **°C/min**

   b. Pressure change over time: **kPa/s or mmHg/s**

   c. Mass change over time: **g of C/min**

2. Define reaction rate.

   The rate of a reaction refers to the speed at which a chemical reaction occurs.

Learning Activity 3.2: Average and Instantaneous Rates

1. Explain how average and instantaneous rates are calculated.

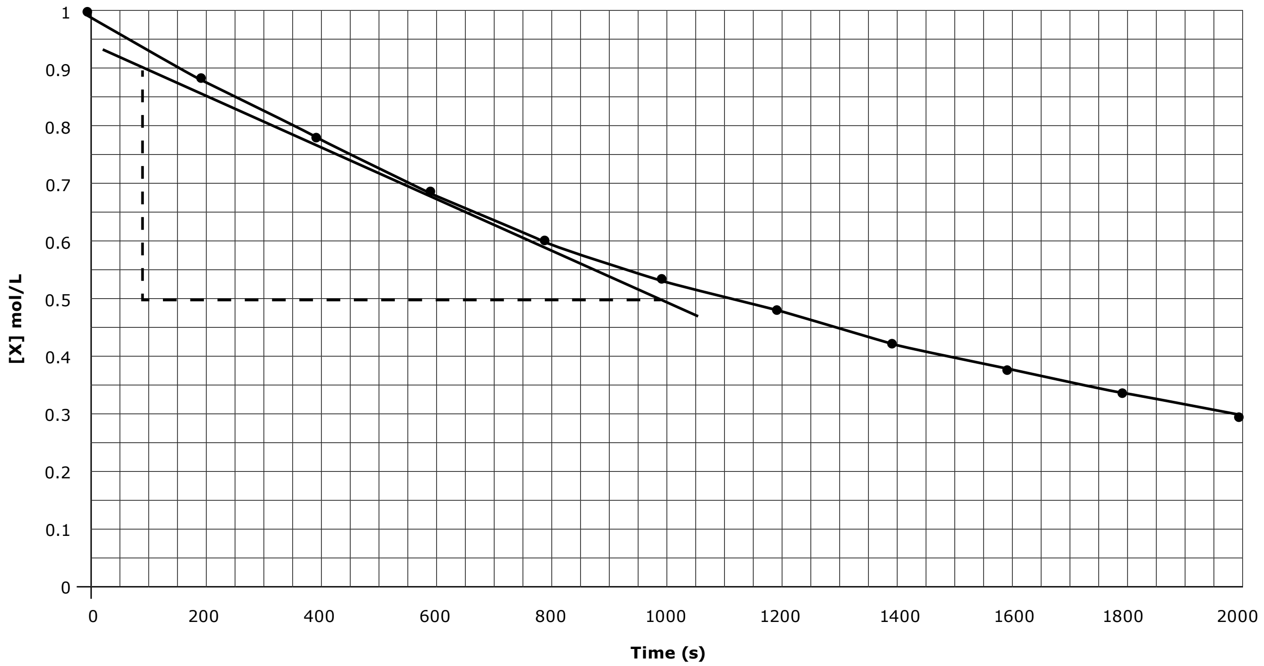
   The average rate of a chemical reaction is calculated by taking an interval of change in data (mass, concentration, volume) and comparing it (dividing) by the time change. The instantaneous rate is calculated by drawing a tangent to the graphed data and determining the slope of that line to indicate the rate of change for that instant.

2. Explain the differences between average and instantaneous rates.

   The average rate is representative of the entire set of data and measures the change in a measured value for a large time interval. The instantaneous rate is representative of the change in a measured value for a particular instant.

3. Graph the following data. Plot *time* on the x-axis and *concentration* on the y-axis.

Time (*s*)	[X] mol/L
0	1.00
200	0.88
400	0.78
600	0.69
800	0.61
1000	0.54
1200	0.48
1400	0.43
1600	0.38
1800	0.34
2000	0.30



4. a. Explain how you would calculate the average rate of reaction for the decomposition of X over time.

   To calculate the average rate of reaction, you could take a reading of [X] at an instant in time (t_2) and a second reading of [X] at another instant in time (t_1). You could then take the difference between these two concentration values and divide by the interval of time ($t_2 - t_1$).

   b. Calculate the average rate of the reaction in the first 200 seconds.

   (0.88 mol/L – 1.00 mol/L) / (200s – 0s) = 6.0×10^{-4} mol/L/s

   c. Calculate the average rate of the reaction in the last 200 seconds.

   (0.30 mol/L – 0.34 mol/L) / (2000s – 1800s) = 2.0×10^{-4} mol/L/s

   d. Draw a tangent at t = 400s. Calculate the instantaneous rate for this time.

   Rate = (0.4 mol/L) / 850 s = 4.70×10^{-4} mol/L/s

Learning Activity 3.3: The Collision Theory

1. Complete the following statements.

   a. The collision theory states that ___**molecules/atoms/particles**___ must collide in order to react.

   b. An effective collision leads to the formation of ___**products**___.

   c. In a(n) ___**effective collision**___, particles must be oriented in a favourable position that allows the bonds of reactants to break.

   d. The energy required to break bonds between reactants comes from the ___**kinetic energy**___ of reactant particles.

   e. When particles collide with energy equal to the activation energy, existing ___**bonds**___ are disrupted.

Learning Activity 3.4: Factors Affecting Reaction Rate

1. In general, what effect does an increase in the concentration of the reactants have on the rate of the reaction? Explain the effect using the collision theory.

   Increasing the concentration of reactants increases the number of particles in the container. An increase in the total number of particles increases the number of particles with kinetic energies greater than or equal to the activation energy.

   More particles having the required activation energy results in more frequent, effective collisions.

   More effective collisions result in an increased reaction rate.

   Increasing the number of particles in a container reduces the space between the particles.

   Smaller spaces between the particles result in a greater probability of collisions occurring and, therefore, an increased frequency of collisions.

2. How do changes in each of the following factors affect the rate of a chemical reaction?

   a. Temperature

      Increasing the temperature increases the rate of most reactions. Decreasing the temperature decreases the rate. Increasing the temperature increases the kinetic energy of all the particles; therefore, more particles have the required activation energy to increase the reaction rate. Decreasing the temperature decreases the kinetic energy of all the particles. If there are fewer particles with the required activation energy, the rate will decrease.

   b. Particle size

      Grinding the reactants increases the surface area and the number of reactant particles in contact. The increased number of particles in contact increases the frequency of collisions and the reaction rate.

   c. Pressure

      Pressure changes only affect gaseous reactants. Increasing the pressure reduces the space between particles, which increases the probability of effective collisions. An increased probability of collisions increases their frequency and rate. Decreasing the pressure increases the space between particles, which decreases the frequency of collisions and, therefore, the rate.

3. List two ways to increase the rate of the following reaction:

   $$Ca_{(s)} + 2\,H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

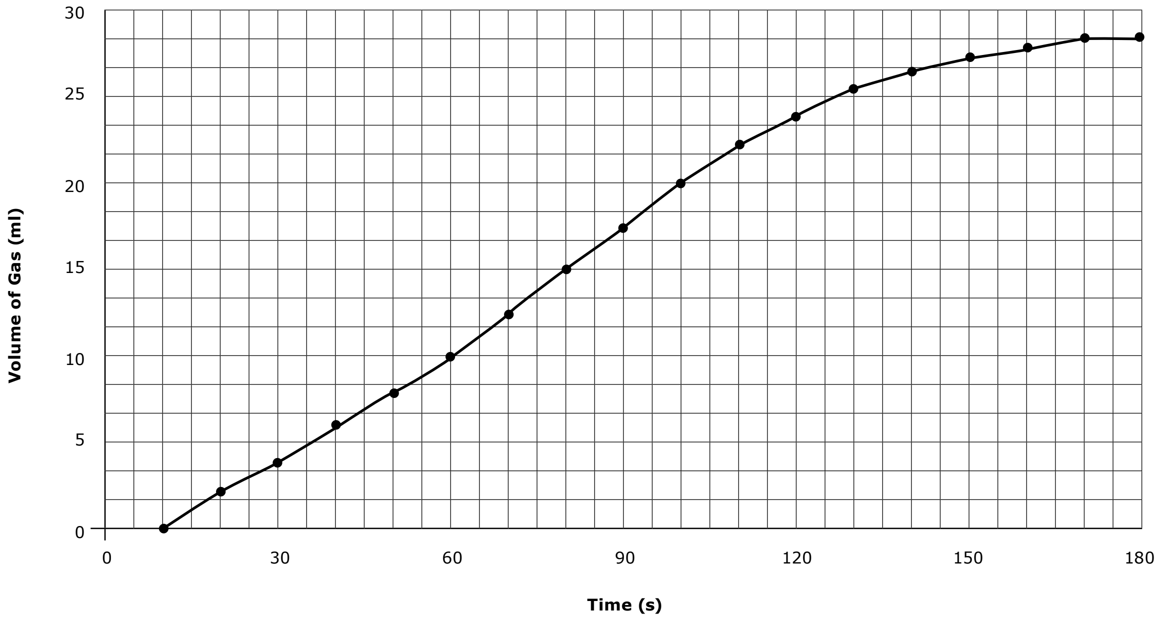
   For each, use your knowledge of collision theory to explain why the rate increases.

   - **Increasing the temperature will increase the rate of collisions between particles of reactants.**
   - **Increasing the surface area of the solid Ca will result in more particles in contact with each other and will, therefore, increase the reaction rate.**

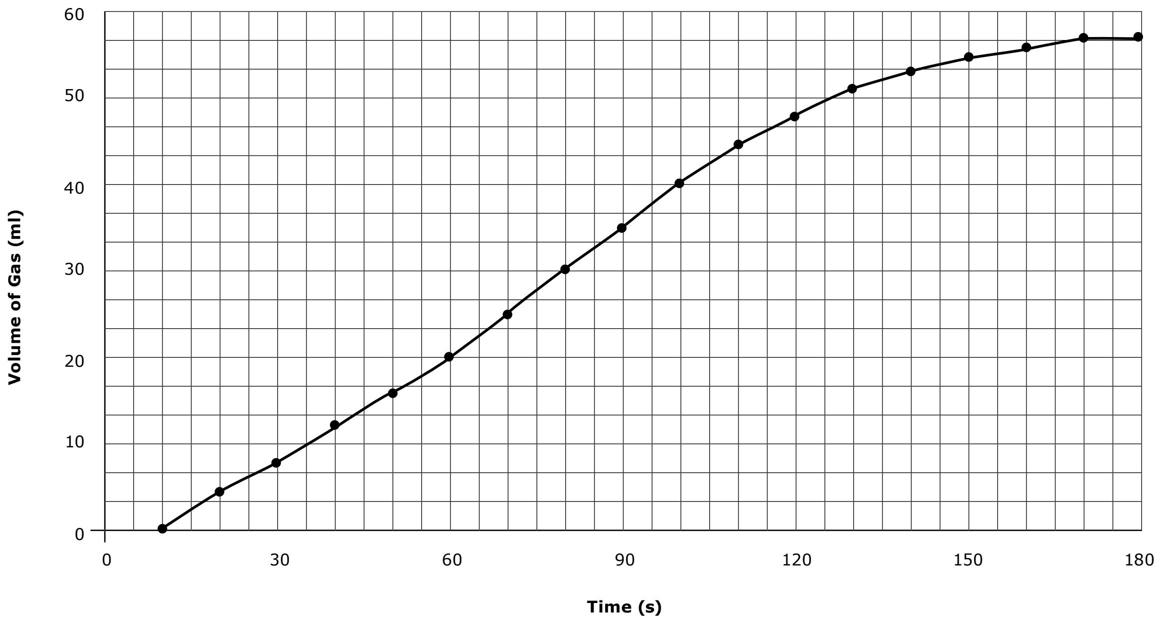
Learning Activity 3.5: Reaction Rate Laboratory Activity

1. Create a graph for each set of data provided **or** use the data you collected to create a graph.

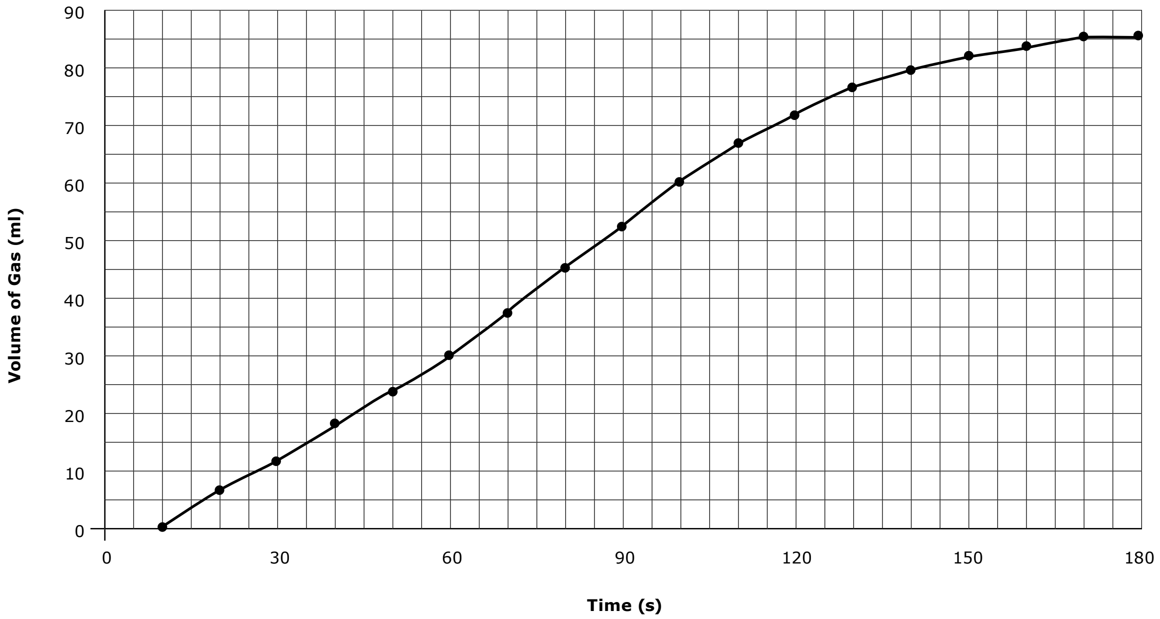
Volume of Gas Collected versus Time (Trials 1 and 3)



Volume of Gas Collected versus Time (Trial 2)



Volume of Gas Collected versus Time (Trial 4)



2. Work with Part A data to answer the following questions:
    a. What is the average rate for Trial 1 (20ºC and one entire tablet)?

        27.9 mL/ 180 s = 0.155 mL/s

    b. What is the average rate for Trial 2 (30ºC and one entire tablet)?

        55.8 mL/ 180 s = 0.31 mL/s

    c. What effect did the increase in temperature by 10ºC have on the amount of gas collected?

        The amount of gas collected doubled in the same period of time.

    d. What effect did the increase in temperature by 10ºC have on the reaction rate of gas collected?

        The increase in temperature doubled the reaction rate.

    e. Using collision theory, explain your results.

        By increasing the temperature, there are more frequent and effective collisions. By allowing more collisions to occur and having more effective collisions, the rate of reaction increases.

3. Work with part B data for question 3.
    a. What is the average rate for trial 3 (20ºC and one entire tablet)?

        27.9 mL/ 180 s = 0.155 mL/s

    b. What is the average rate for trial 4 (20ºC and one entire tablet crushed into a fine powder)?

        83.7 mL/ 180 s = 0.485 mL/s

    c. What factor was manipulated by crushing the tablet into a fine powder?

        The surface area of the effervescent tablet was increased by crushing it into a fine powder.

    d. What effect did crushing the tablet into a fine powder have on the amount of gas collected?

        The amount of gas collected tripled in the same period of time.

    e. What effect did crushing the tablet into a fine powder have on the reaction rate of gas collected?

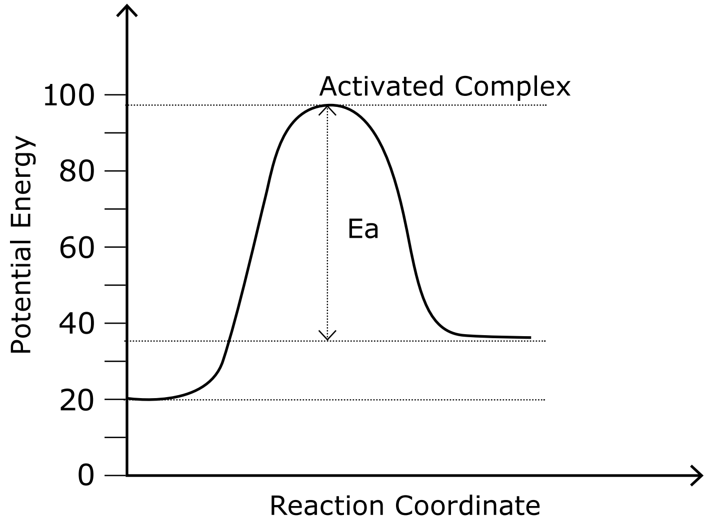
        The increase in surface area tripled the reaction rate when compared to the original data.

    f. Using the collision theory, explain your results.

        By crushing the tablet to dust, the particles have more exposed surface area and are able to react more readily.

Learning Activity 3.6: Potential Energy Diagrams

1. Given the following reaction coordinate diagram:



   a. What is the activation energy of the reaction shown by the diagram above?

   Ea = 95 kJ – 20 kJ = 75 kJ

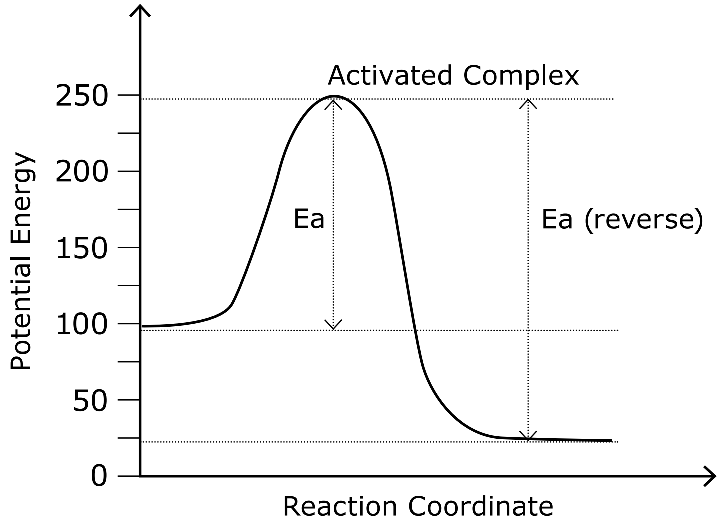
   b. What is the enthalpy change for this reaction?

   $\Delta H = H_{product} - H_{reactant} = 35\ kJ - 20 kJ = 15\ kJ$

   c. Is this reaction endothermic or exothermic?

   Since ΔH is positive, the reaction is endothermic.

2. Given the following reaction coordinate diagram:



a. What is the activation energy of the reaction in the diagram above?

Ea = 250 kJ – 100 kJ = 150 kJ

b. What is the enthalpy change for this reaction?

$\Delta H = H_{product} - H_{reactant} = 25\ kJ - 100\ kJ = -75\ kJ$

c. Is this reaction endothermic or exothermic?

Since ΔH is negative, the reaction is exothermic.

d. What would be the activation energy of the *reverse* reaction?

For the reverse reaction, products now become the reactants; therefore, Ea = 250 kJ – 25 kJ = 225 kJ.

3. What is the activated complex or transition state and how is it related to reaction rates?

The activated complex is a short-lived combination of atoms formed at the top of a potential energy diagram from reactants to product. Lowering the required activation energy of a reaction increases the number of particles with enough energy to produce an effective collision. As a result, the reaction rate will increase.

4. Why does an exothermic reaction produce heat even though its energy value is negative?

Exothermic reactions release energy. This means that energy is lost from the system. Since energy is lost, the system has less energy than it began with (enthalpy of products less than enthalpy of reactants). This produces a negative value.

5. Label each of these reactions as endothermic or exothermic.

a. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat **exothermic**

b. $6C + 3H_2 \rightarrow C_6H_6$, $\Delta H = +83$ kJ **endothermic**

c. $2NO_2 + 112\ kJ \rightarrow 2NO + O_2$ **endothermic**

d. $S + O_2 \rightarrow SO_2$, $\Delta H = -296$ kJ **exothermic**

Learning Activity 3.7: Describing Reaction Mechanism

Answer each of the following in the space provided.

1. Consider the following reaction:
   $X_{(s)} + Y_{(g)} + \text{heat} \rightarrow Z_{(aq)}$
   Suggest ways to increase the reaction rate of this chemical reaction. List all possibilities.
   - **crush X to increase the surface area**
   - **increase Y**
   - **decrease the volume of the container**
   - **increase the pressure on the system**
   - **increase the temperature of the system**
   - **add a catalyst**
2. For each of the reaction mechanisms below
   - Identify the intermediates.
   - Identify the rate determining step.
   - Determine the overall equation.

   a. $H_2 + ICl \rightarrow HI + HCl$ (slow)
   $HI + ICl \rightarrow I_2 + HCl$
   - **The intermediate is HI**
   - **The rate determining step is the slow step:**
     $H_2 + ICl \rightarrow HI + HCl$
   - **The overall equation is $H_2 + 2\,ICl \rightarrow I_2 + 2\,HCl$**

   b. $O_3 \rightarrow O_2 + O$ (slow)
   $O_3 + O \rightarrow 2\,O_2$
   - **The intermediate is: O**
   - **The rate determining step is the slow step: $O_3 \rightarrow O_2 + O$**
   - **The overall equation is $2\,O_3 \rightarrow 3\,O_2$**

3. Describe where each of these would appear in a reaction mechanism:

   Reactants

   Substances which first appear on the LHS of the reaction mechanism and are also on the LHS of the net reaction.

   Products

   Substances which first appear on the RHS of the reaction mechanism and are also on the RHS of the net reaction.

   Intermediates

   For substances which first appear on the RHS of the reaction mechanism and then appear on the LHS in the following step of the mechanism, the substance is cancelled out and does not appear in the net reaction.

   Catalysts

   For substances which first appear on the LHS of the reaction mechanism and then appear on the RHS in the following step of the mechanism, the substance is cancelled out and does not appear in the net reaction.

4. Describe a reaction mechanism.

   In any chemical change, some bonds are broken and new ones are formed. Quite often, these changes are too complicated to happen in one simple step. Instead, the reaction may involve a series of small intermediate reactions, one after the other. A reaction mechanism describes the one or more steps involved in the reaction *(1 mark)* in a way that clearly shows exactly how the various bonds are broken and made. *(1 mark)*

5. How are the elementary steps in a reaction related to the overall reaction mechanism?

   The overall mechanism is the sum of all the elementary steps.

Learning Activity 3.8: Understanding Rate Law

1. For the reaction A → products, the following data was collected:

Time (min)	Mass of A (g)
0.0	25.0
1.0	20.0
2.0	17.0
3.0	15.0
4.0	13.0
5.0	12.0

   a. Calculate the average rate, in g A/min, over the entire 5 minutes.

   13g / 5.0 min = 2.6 g/min

   b. Calculate the average rate for the interval between 2.0 and 4.0 minutes.

   4.0 g / 2.0 min = 2.0 g/min

   c. Calculate the average rate for the interval between 3.0 and 4.0 minutes.

   2.0 g / 1.0 min = 2.0 g/min

2. The decomposition of acetaldehyde to methane and carbon dioxide occurs according to the following equation:

$$CH_3CHO_{(g)} \rightarrow CH_{4(g)} + CO_{(g)}$$

The results of an experiment are given below:

Time (s)	$[CH_3CHO]$ (mol/L)
42	0.00667
73	0.00626
105	0.00586
190	0.00505
242	0.00464
310	0.00423
384	0.00383
480	0.00342
665	0.00282
840	0.00241

2. a. Calculate the rate of decomposition of acetaldehyde between 42 s and 105 s.

   (–0.00081 mol/L) / 63 s = 1.3 x 10^{-5} mol/L/s

   Note that rate is always expressed as a positive value.

   b. Calculate the rate of decomposition in the interval 190 s to 480 s.

   (–0.00163 mol/L) / 290 s = 5.6 x 10^{-6} mol/L/s

   Note that rate is always expressed as a positive value.

3. A first-order reaction initially proceeds at a rate of 0.5 mol/L·s. What is the rate when *half* the starting material remains? When *a quarter* of the starting material remains?

   A first-order reaction means the concentration change is proportional to the rate change in a ratio of 1:1. When half the starting material remains, the rate will be 0.25 mol/L·s. When a quarter of the starting material remains, the rate will be 0.125 mol/L·s.

4. How does the magnitude of the k value relate to the speed of the reaction?

   If the k value if large, the products form quickly. If the k value is low, the products form slowly.

5. The rate law for the following reaction is second order and first order with respect to NO. Write the rate law expression for this reaction.

   $NO_{(g)} + O_{3(g)} \rightarrow NO_{2(g)} + O_{2(g)}$

   Rate = $K[NO]^1[O_3]^1$ or Rate = $K[NO][O_3]$

6. The rate law for the reaction of nitrogen-dioxide and fluorine gases is as follows:

   Rate = $K[NO_2]^1[F_2]^1$

   What will happen to the reaction rate if the F_2 concentration is tripled? Show your work.

   The reaction rate will increase by a factor of three.

   Example of rationale

   Rate = $K[1][1]$ = 1

   Rate = $K[1][3]$ = 3

   The rate increased by a factor of three.

7. Use the equation and data table below to answer the following questions.

$NO_{(g)} + O_{2(g)} \rightarrow NO_{2(g)}$

Experiment	Initial $[O_2]$ (mol/L)	Initial [NO] (mol/L)	Initial Reaction Rate (mol/dm^3·s)
1	0.0010	0.0010	7.10
2	0.0040	0.0010	28.4
3	0.0040	0.0030	256
4	0.0020	0.0010	14.1

a. Determine the rate law for the reaction.

Order with respect to O_2 is $4^1 = 4$, first order (using experiments 1 and 2)

Order with respect to NO is $3^2 = 9$, second order (using experiments 2 and 3)

Rate $= K[NO]^2[O_2]^1$

b. Determine the overall order of the reaction.

2 + 1 = third order overall

GRADE 12 CHEMISTRY

Midterm Practice Examination

Name: ____

Student Number: ______________________________

Attending ☐ Non-Attending ☐

Phone Number: ______________________________

Address: ____________________________________

____

____

For Marker's Use Only

Date: ______________________________

Final Mark _______ /100 = ________%

Comments:

Instructions

You have a maximum of 3 hours to complete your final examination.
The final examination will be weighted as follows:

- Modules 1 to 3: 100%

The format of the examination will be as follows:

- Part A: Multiple Choice *(27 x 1 = 27 marks)*
- Part B: Fill-in-the-Blanks *(23 x 1 = 23 marks)*
- Part C: Short Answer *(50 marks)*

Include units with all answers as required.

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- scrap paper
- a ruler
- a graphing or scientific calculator

The following resources are provided at the end of this examination:

- Periodic Table of Elements
- Alphabetical Listing of the Elements and Their Atomic Masses
- Table of Electronegativities
- Relative Strengths of Acids Table
- Solubility Chart
- Table of Standard Reduction Potentials with Values
- Names, Formulas, and Charges of Common Ions
- Common Ions

Part A: Multiple Choice *(27 marks total)*

Use the answer sheet found at the end of this examination to answer the multiple-choice questions in this section. Shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Aqueous Reactions *(7 marks)*

1. Oxidation is defined as a
   a. gain of electrons
   b. gain of hydrogen
   c. loss of electrons
   d. loss of oxygen
2. Which of the following statements about an oxidation number is FALSE?
   a. The number can be positive.
   b. The number can be assigned to a molecule.
   c. The number can be negative.
   d. The numbers are assigned according to a set of rules.
3. As a free element, the oxidation number of chlorine is
   a. +1
   b. –1
   c. +2
   d. 0
4. In which reaction is oxidation occurring?
   a. $Mg^{2+} + 2e^- \rightarrow Mg$
   b. $Ba^{2+} + S^{2-} \rightarrow BaS$
   c. $NH_3 + H^+ \rightarrow NH_4^+$
   d. $2Cl^- \rightarrow Cl_2 + 2e^-$
5. Identify the reducing agent in the following reaction:
   $Na_{(s)} + H_20_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$
   a. Na
   b. H_2O
   c. NaOH
   d. H_2

Name: ____

6. Determine what happens in the following reaction:
   $S + Cl_2 \rightarrow SCl_2$
   a. Chlorine is oxidized.
   b. Chlorine is reduced.
   c. Sulphur is reduced.
   d. Sulfur is the oxidizing agent.

7. Which of these is **not** a strong acid?
   a. hydrochloric acid (HCl)
   b. hydrobromic acid (HBr)
   c. sulfuric acid (H_2SO_4)
   d. acetic acid (CH_3COOH)

Atomic Structure *(10 marks)*

8. The electron configuration for calcium is
   a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 5p^4$
   b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
   c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
   d. none of the above

9. Which of the following elements has the HIGHEST first ionization energy?
   a. N
   b. As
   c. P
   d. Bi

10. Which of the following elements is the LEAST electronegative?
    a. Cl
    b. K
    c. I
    d. Ge

11. The number of valence electrons for the selenium atom is
    a. 34
    b. 16
    c. 4
    d. 6

12. Which atom has the smallest atomic radius?
    a. Mg
    b. Al
    c. Si
    d. S

13. What is the highest principal quantum number (*n*), according to the quantum mechanical model?
    a. 2
    b. 7
    c. 4
    d. 1

14. How many electrons can be accommodated in the third principle energy level?
    a. 2
    b. 8
    c. 18
    d. 34

15. Write the complete electronic configuration for the Zn^{2+} ion.
    a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^{10}$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
    c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
    d. $[Ar] 4s^0 3d^{10}$

16. The quantum mechanical model of the atom
    a. defines the exact path of an electron around the nucleus
    b. was proposed by Neils Bohr
    c. defines the propability of finding an electron in a certain position
    d. states that electrons can be found in circular orbitals around the nucleus

17. Stable electron configurations are likely to contain
    a. filled energy sublevels
    b. unfilled s-orbitals
    c. electrons with a clockwise spin
    d. filled d sublevels

Name: ______________________________________

Kinetics ***(10 marks)***

18. Which of these is the reason that a reaction will speed up when there is an increase in temperature?
    a. Chemical bonds are more likely to break.
    b. Collisions between reactant molecules are more accurate.
    c. There are more collisions with greater energy.
    d. The activation energy increases as temperature increases.

19. Which expression represents a reaction rate?
    a. $\frac{\text{time}}{\text{mass}}$
    b. $\frac{\text{energy}}{\text{mass}}$
    c. $\frac{\text{mass}}{\text{time}}$
    d. $\frac{\text{time}}{\text{energy}}$

20. The rate of a chemical reaction usually
    a. increases as reactant concentration increases
    b. decreases as reactant concentration increases
    c. decreases as temperature increases
    d. is slowed down by a catalyst

21. If a catalyst is used in a reaction,
    a. the activation energy will increase
    b. the reaction rate increases
    c. different products will be made
    d. the reaction rate decreases

22. Assuming it occurs in a single step, what is the rate law for the following reaction? $A + 2B \rightarrow C + D$
    a. Rate = $K[A][B]$
    b. Rate = $K[A]2[B]$
    c. Rate = $K[A][B]^2$
    d. Rate = $K[A][B]^{1/2}$

23. Given a first order reaction where the concentration of the reactants is doubled, which of the following is true?
    a. The rate does not change.
    b. The rate doubles.
    c. The rate quadruples.
    d. The rate decreases.

24. An elementary reaction
    a. has only elements as reactants
    b. has one single step
    c. has only elements as products
    d. never needs a catalyst

25. What information is **not** given by an overall balanced chemical equation?
    a. the reaction mechanism
    b. the number of atoms participating in the reaction
    c. the probable order of the reaction
    d. the relative number of molecules used in the reaction

26. Given the following reaction and assuming a single step,
    $2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)}$ $\Delta H = -91$ kcal, if the volume of the enclosing vessel is halved, the reaction rate will _______________.
    a. double
    b. quadruple
    c. increase eight-fold
    d. decrease by half

27. For the rate law expression, rate = $K[A]^2[B][C]^2$, the reaction is ___________ order overall.
    a. second
    b. third
    c. fifth
    d. seventh

Name: ____

Part B: Fill-in-the-Blank *(23 marks total)*

Use the answer sheet found at the end of this examination to answer the fill-in-the-blank questions of this section. Write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Using a term from the word bank provided below, complete each of the statements that follow. Each blank is worth one mark; therefore, some questions have a total value of two marks. There are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use.

4
absorbed
activation energy
amplitude
Aufbau principle
collision theory
concentration
configuration
d
decrease
dissociate
dissolution
effective
effective collision
energy level
excited
f
frequency
gain
ground
Heisenberg uncertainty
Hund's rule
increase
kinetic
kinetic energy
loss
mechanism
molecules
motion
n
neutralization
nine
opposite
orbital(s)
order
orientation
oxidation
oxidation number
oxidizing agent
p
Pauli exclusion
peak
photon(s)
precipitation
probability
product(s)
rate
rate law
reactant(s)
reaction mechanism
reaction rate
reducing agent
reduction
s
salt
Schrödinger
separate
solubility
soluble
solvation
speed
sublevels
titration
valence
water
wavelike

Aqueous Reactions ***(5 marks)***

1. A substance is considered _____________ if it will dissolve in a specific solvent.
2. An _____________ in the oxidation number of an atom signifies oxidation, while a _____________ in the oxidation number signifies reduction.
3. A _____________ reaction is one in which the aqueous (dissolved) ions combine to form a solid that is not readily soluble.
4. _____________ is the addition of a known amount of solution to determine the volume (or concentration) of another solution.

Atomic Structure ***(10 marks)***

5. The _____________ orbitals are only present in the fourth energy level and up.
6. Energy is _____________ when an electron moves from ground state to excited state.
7. Electron orbitals are areas of high _____________ of finding electrons.
8. Pauli said that two electrons can only occupy the same orbital if they have _____________ spins.
9. The electrons on the highest numbered subshells are the _____________ electrons.
10. The _____________ _____________ states that electrons must be added to the lowest energy state before adding electrons to a higher energy level.
11. The _____________ of a wave is the height of the wave from the origin to the top of the crest.
12. Einstein proposed that light is composed of particles, or quanta, known as _____________.
13. The quantum mechanical model of the atom was proposed by _____________.

Kinetics ***(8 marks)***

14. The _____________ of a chemical reaction can be determined from the balanced chemical equation.
15. A reaction mechanism describes the _____________ in which bonds break and atoms rearrange in a chemical reaction.
16. If changing the concentration of a product does not change the rate, it does not appear in the _____________ _____________.
17. In an _____________ _____________, the particles are oriented in a favourable position that allows bonds to break and atoms to rearrange.
18. The difference between the energy at the peak of the activation diagram and the energy of the reactants is called the _____________ _____________.

Name: ____

Part C: Short Answer *(50 marks total)*

Answer each of the following questions 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 as well as the appropriate unit.

Aqueous Reactions *(25 marks)*

1. Give the oxidation number for each atom in Na_3AlF_6. Show your work. *(4 marks)*

2. Complete the dissociation equation for the following ionic compound. *(2 marks)*

   $NaF_{(s)} \rightarrow$

3. Balance the following redox reaction using the oxidation-number change method. You will earn up to 5 marks for the following steps:
   - 1 mark: assigning correct oxidation numbers
   - 1 mark: showing electron loss
   - 1 mark: showing electron gain
   - 1 mark: assigning correct coefficients
   - 1 mark: simplifying the terms that appear on both sides of the equation

   $Al_{(s)} + MnO_{2(s)} \rightarrow Al_2O_{3(s)} + Mn_{(s)}$

4. Why must oxidation and reduction occur together? *(2 marks)*

Name: ____

5. If 40.8 mL of 0.106 mol/L sulphuric acid neutralizes 61.8 mL of potassium hydroxide solution, answer the following questions to determine the concentration of the base.
   a. Write the balanced chemical equation for the neutralization reaction. *(2 marks)*

   b. Use the concentration and volume of the known acid or base to calculate the moles of the known substance (acid or base). *(2 marks)*

   c. Use the coefficients from the balanced equation to determine the moles of the unknown substance (acid or base). *(2 marks)*

d. Calculate the required concentration of the base. *(2 marks)*

6. Write the balanced neutralization reaction given the following reactants. *(2 marks)*

   $2\ HCl + Mg(OH)_2 \rightarrow$

7. Use the *solubility rules* (see page 21) to predict whether a precipitate will form when the following pair of aqueous solutions are mixed. If a precipitate forms, write the balanced net ionic equation for the reaction. *(2 marks)*

   $Pb(C_2H_3O_2)_{2(aq)}$ and $LiCl_{(aq)}$

Name: ____

Atomic Structure ***(8 marks)***

8. Write the electron configuration and noble gas configuration for a chlorine atom. *(2 marks)*

9. Outline the contributions of Neils Bohr, Werner Heisenberg, and Erwin Schrödinger in relation to our understanding of the quantum mechanical model of the atom. *(3 marks)*

10. Write the electronic configurations for the valence electrons of each of the following. *(3 marks)*

    a. Mg^{2+}

    b. S^{2-}

    c. sodium

Kinetics ***(17 marks)***

11. For the reaction $CH_3COCH_3 + I_2 \rightarrow CH_3COCH_2I + HI$, the following data was obtained:

Experiment	Initial $[CH_3\ COCH_3]$	Initial $[I_2]$	Reaction Rate (mol/L · s)
1	0.100 mol/L	0.100 mol/L	1.16×10^{-7}
2	0.0500 mol/L	0.100 mol/L	5.18×10^{-8}
3	0.0500 mol/L	0.500 mol/L	5.18×10^{-8}

    a. Write the rate expression for the reaction. *(1 mark)*

    b. What order is the reaction with respect to each of the reactants? *(2 marks)*

Name: ____

c. What is the overall order of this reaction? *(1 mark)*

d. Calculate the specific rate constant. *(2 marks)*

e. What is the reaction rate if the initial concentrations are both 0.045 mol/L? *(2 marks)*

12. List two ways in which reaction rates can generally be increased. *(2 marks)*

13. Consider the following reaction mechanism.

$H_2O_2 + I^- \rightarrow H_2O + IO^-$ (slow) $\Delta H = 45$ kJ
$H_2O_2 + IO^- \rightarrow H_2O + O_2 + I^-$ (fast) $\Delta H = 35$ kJ

a. What is (are) the intermediate(s)? *(1 mark)*

-

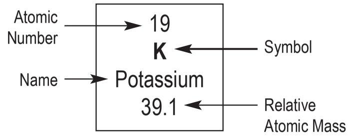
b. What is the overall equation? *(2 marks)*

c. Draw a potential energy diagram for the reaction. *(3 marks)*

d. In which step would a catalyst be most useful? *(1 mark)*

Name: ______________________________

Periodic Table of Elements



	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)

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)	Irdium	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)	Sulphur	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
Einstienium	(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		

Name: ______________________________________

Table of Electronegativities

	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** —

Relative Strengths of Acids Table

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulphuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4×10^{-2}
Sulphurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7×10^{-2}
Hydrogen sulphate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3×10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1×10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0×10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3×10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7×10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1×10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7×10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5×10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6×10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4×10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4×10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4×10^{-7}
Hydrogen sulphide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3×10^{-8}
Hydrogen sulphite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2×10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7×10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7×10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0×10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4×10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4×10^{-13}
Hydrogen sulphide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2×10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8×10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Name: ____

Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	low solubility
	all others	soluble
sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulphide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulphite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

Table of Standard Reduction Potentials with Values

Oxidized Species	↔	Reduced Species	E°/V
$Li^{+}_{(aq)} + e^{-}$	↔	$Li_{(s)}$	–3.04
$K^{+}_{(aq)} + e^{-}$	↔	$K_{(s)}$	–2.93
$Ca^{2+}_{(aq)} + 2e^{-}$	↔	$Ca_{(s)}$	–2.87
$Na^{+}_{(aq)} + e^{-}$	↔	$Na_{(s)}$	–2.71
$Mg^{2+}_{(aq)} + 2e^{-}$	↔	$Mg_{(s)}$	–2.37
$Al^{3+}_{(aq)} + 3e^{-}$	↔	$Al_{(s)}$	–1.66
$Mn^{2+}_{(aq)} + 2e^{-}$	↔	$Mn_{(s)}$	–1.19
$H_2O_{(l)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)} + OH^{-}_{(aq)}$	–0.83
$Zn^{2+}_{(aq)} + 2e^{-}$	↔	$Zn_{(s)}$	–0.76
$Fe^{2+}_{(aq)} + 2e^{-}$	↔	$Fe_{(s)}$	–0.45
$Ni^{2+}_{(aq)} + 2e^{-}$	↔	$Ni_{(s)}$	–0.26
$Sn^{2+}_{(aq)} + 2e^{-}$	↔	$Sn_{(s)}$	–0.14
$Pb^{2+}_{(aq)} + 2e^{-}$	↔	$Pb_{(s)}$	–0.13
$H^{+}_{(aq)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}_{(aq)} + e^{-}$	↔	$Cu^{+}_{(aq)}$	+0.15
$SO_4^{2-}{}_{(aq)} + 4\ H^{+}_{(aq)} + 2e^{-}$	↔	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}_{(aq)} + 2e^{-}$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^{-}$		$2OH^{-}_{(aq)}$	+0.40
$Cu^{+}_{(aq)} + e^{-}$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^{-}$	↔	$I^{-}_{(aq)}$	+0.54
$Fe^{3+}_{(aq)} + e^{-}$	↔	$Fe^{2+}_{(aq)}$	+0.77
$Ag^{+}_{(aq)} + e^{-}$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^{-}$	↔	$Br^{-}_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^{+}_{(aq)} + 2e^{-}$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7^{2-}{}_{(aq)} + 14H^{+}_{(aq)} + 6e^{-}$	↔	$2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^{-}$	↔	$Cl^{-}_{(aq)}$	+1.36
$MnO_4^{-}{}_{(aq)} + 8H^{+}_{(aq)} + 5e^{-}$	↔	$Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^{-}$	↔	$F^{-}_{(aq)}$	+2.87

Name: ____

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

Names, Formulas, and Charges of Common Ions *(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^-	sulphate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulphate	HSO_4^-
fluoride	F^-	sulphide	S^{2-}
hydroxide	OH^-	hydrogen sulphide	HS^-
hypochlorite	ClO^-	sulphite	SO_3^{2-}
iodide	I^-	hydrogen sulphite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Name: ____

Common Ions

Cations (Positive Ions)

1^{+} charge		2^{+} charge		3^{+} charge	
$NH4^{+}$	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	$Pb2^{2+}$	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

Common Ions *(continued)*

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulphide	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-}	Sulphate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulphide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulphite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulphate
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 sulphate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Name: ____

Midterm Practise Examination Answer Sheet

Part A: Multiple Choice *(27 marks)*

For each multiple-choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Example: Ⓐ Ⓑ ● Ⓓ

1. Ⓐ Ⓑ Ⓒ Ⓓ
2. Ⓐ Ⓑ Ⓒ Ⓓ
3. Ⓐ Ⓑ Ⓒ Ⓓ
4. Ⓐ Ⓑ Ⓒ Ⓓ
5. Ⓐ Ⓑ Ⓒ Ⓓ
6. Ⓐ Ⓑ Ⓒ Ⓓ
7. Ⓐ Ⓑ Ⓒ Ⓓ
8. Ⓐ Ⓑ Ⓒ Ⓓ
9. Ⓐ Ⓑ Ⓒ Ⓓ
10. Ⓐ Ⓑ Ⓒ Ⓓ
11. Ⓐ Ⓑ Ⓒ Ⓓ
12. Ⓐ Ⓑ Ⓒ Ⓓ
13. Ⓐ Ⓑ Ⓒ Ⓓ
14. Ⓐ Ⓑ Ⓒ Ⓓ
15. Ⓐ Ⓑ Ⓒ Ⓓ
16. Ⓐ Ⓑ Ⓒ Ⓓ
17. Ⓐ Ⓑ Ⓒ Ⓓ
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20. Ⓐ Ⓑ Ⓒ Ⓓ
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22. Ⓐ Ⓑ Ⓒ Ⓓ
23. Ⓐ Ⓑ Ⓒ Ⓓ
24. Ⓐ Ⓑ Ⓒ Ⓓ
25. Ⓐ Ⓑ Ⓒ Ⓓ
26. Ⓐ Ⓑ Ⓒ Ⓓ
27. Ⓐ Ⓑ Ⓒ Ⓓ

Part B: Fill-in-the-Blank *(23 marks)*

For each fill-in-the-blank question, write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Aqueous Reactions ***(5 marks)***

1. ____
2. ____

   ____
3. ____
4. ____

continued

Atomic Structure *(10 marks)*

5. ____

6. ____

7. ____

8. ____

9. ____

10. ____

____

11. ____

12. ____

13. ____

Kinetics *(8 marks)*

14. ____

15. ____

16. ____

____

17. ____

____

18. ____

____

Grade 12 Chemistry

Midterm Practice Examination
Answer Key

Name: ____

Student Number: ______________________________

Attending ☐ Non-Attending ☐

Phone Number: ________________________________

Address: _____________________________________

____

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For Marker's Use Only

Date: ____________________________

Final Mark ______/100 = ________%

Comments:

Instructions

You have a maximum of 3 hours to complete your final examination.
The final examination will be weighted as follows:

- Modules 1 to 3: 100%

The format of the examination will be as follows:

- Part A: Multiple Choice *(27 x 1 = 27 marks)*
- Part B: Fill-in-the-Blanks *(23 x 1 = 23 marks)*
- Part C: Short Answer *(50 marks)*

Include units with all answers as required.

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- scrap paper
- a ruler
- a graphing or scientific calculator

The following resources are provided at the end of this examination:

- Periodic Table of Elements
- Alphabetical Listing of the Elements and Their Atomic Masses
- Table of Electronegativities
- Relative Strengths of Acids Table
- Solubility Chart
- Table of Standard Reduction Potentials with Values
- Names, Formulas, and Charges of Common Ions
- Common Ions

Part A: Multiple Choice *(27 marks total)*

Use the answer sheet found at the end of this examination to answer the multiple-choice questions in this section. Shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Aqueous Reactions *(7 marks)*

1. Oxidation is defined as a
   a. gain of electrons
   b. gain of hydrogen
   c. loss of electrons
   d. loss of oxygen
2. Which of the following statements about an oxidation number is FALSE?
   a. The number can be positive.
   b. The number can be assigned to a molecule.
   c. The number can be negative.
   d. The numbers are assigned according to a set of rules.
3. As a free element, the oxidation number of chlorine is
   a. +1
   b. –1
   c. +2
   d. 0
4. In which reaction is oxidation occurring?
   a. $Mg^{2+} + 2e^- \rightarrow Mg$
   b. $Ba^{2+} + S^{2-} \rightarrow BaS$
   c. $NH_3 + H^+ \rightarrow NH_4^+$
   d. $2Cl^- \rightarrow Cl_2 + 2e^-$
5. Identify the reducing agent in the following reaction:
   $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$
   a. Na
   b. H_2O
   c. NaOH
   d. H_2

Name: ____

6. Determine what happens in the following reaction:
   $S + Cl_2 \rightarrow SCl_2$
   a. Chlorine is oxidized.
   b. Chlorine is reduced.
   c. Sulphur is reduced.
   d. Sulfur is the oxidizing agent.

7. Which of these is **not** a strong acid?
   a. hydrochloric acid (HCl)
   b. hydrobromic acid (HBr)
   c. sulfuric acid (H_2SO_4)
   d. acetic acid (CH_3COOH)

Atomic Structure *(10 marks)*

8. The electron configuration for calcium is
   a. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^25p^4$
   b. $1s^22s^22p^63s^23p^64s^2$
   c. $1s^22s^22p^63s^23p^64s^23d^2$
   d. none of the above

9. Which of the following elements has the HIGHEST first ionization energy?
   a. N
   b. As
   c. P
   d. Bi

10. Which of the following elements is the LEAST electronegative?
    a. Cl
    b. K
    c. I
    d. Ge

11. The number of valence electrons for the selenium atom is
    a. 34
    b. 16
    c. 4
    d. 6

12. Which atom has the smallest atomic radius?
    a. Mg
    b. Al
    c. Si
    d. S

13. What is the highest principal quantum number (n), according to the quantum mechanical model?
    a. 2
    b. 7
    c. 4
    d. 1

14. How many electrons can be accommodated in the third principle energy level?
    a. 2
    b. 8
    c. 18
    d. 34

15. Write the complete electronic configuration for the Zn^{2+} ion.
    a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^{10}$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
    c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
    d. $[Ar]4s^0 3d^{10}$

16. The quantum mechanical model of the atom
    a. defines the exact path of an electron around the nucleus
    b. was proposed by Neils Bohr
    c. defines the propability of finding an electron in a certain position
    d. states that electrons can be found in circular orbitals around the nucleus

17. Stable electron configurations are likely to contain
    a. filled energy sublevels
    b. unfilled s-orbitals
    c. electrons with a clockwise spin
    d. filled d sublevels

Name: ____

Kinetics ***(10 marks)***

18. Which of these is the reason that a reaction will speed up when there is an increase in temperature?
    a. Chemical bonds are more likely to break.
    b. Collisions between reactant molecules are more accurate.
    c. There are more collisions with greater energy.
    d. The activation energy increases as temperature increases.

19. Which expression represents a reaction rate?
    a. $\frac{\text{time}}{\text{mass}}$
    b. $\frac{\text{energy}}{\text{mass}}$
    c. $\frac{\text{mass}}{\text{time}}$
    d. $\frac{\text{time}}{\text{energy}}$

20. The rate of a chemical reaction usually
    a. increases as reactant concentration increases
    b. decreases as reactant concentration increases
    c. decreases as temperature increases
    d. is slowed down by a catalyst

21. If a catalyst is used in a reaction,
    a. the activation energy will increase
    b. the reaction rate increases
    c. different products will be made
    d. the reaction rate decreases

22. Assuming it occurs in a single step, what is the rate law for the following reaction? $A + 2B \rightarrow C + D$
    a. Rate = $K[A][B]$
    b. Rate = $K[A]2[B]$
    c. Rate = $K[A][B]^2$
    d. Rate = $K[A][B]^{½}$

23. Given a first order reaction where the concentration of the reactants is doubled, which of the following is true?
    a. The rate does not change.
    b. The rate doubles.
    c. The rate quadruples.
    d. The rate decreases.

24. An elementary reaction
    a. has only elements as reactants
    b. has one single step
    c. has only elements as products
    d. never needs a catalyst

25. What information is **not** given by an overall balanced chemical equation?
    a. the reaction mechanism
    b. the number of atoms participating in the reaction
    c. the probable order of the reaction
    d. the relative number of molecules used in the reaction

26. Given the following reaction and assuming a single step,
    $2\,SO_{2(g)} + O_{2(g)} \rightarrow 2\,SO_{3(g)}$ $\Delta H = -91$ kcal, if the volume of the enclosing vessel is halved, the reaction rate will _______________.
    a. double
    b. quadruple
    c. increase eight-fold
    d. decrease by half

27. For the rate law expression, rate = $k[A]^2[B][C]^2$, the reaction is ___________ order overall.
    a. second
    b. third
    c. fifth
    d. seventh

Name: ____

Part B: Fill-in-the-Blank *(23 marks total)*

Use the answer sheet found at the end of this examination to answer the fill-in-the-blank questions of this section. Write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Using a term from the word bank provided below, complete each of the statements that follow. Each blank is worth one mark; therefore, some questions have a total value of two marks. There are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use.

4
absorbed
activation energy
amplitude
Aufbau principle
collision theory
concentration
configuration
d
decrease
dissociate
dissolution
effective
effective collision
energy level
excited
f
frequency
gain
ground
Heisenberg uncertainty
Hund's rule
increase
kinetic
kinetic energy
loss
mechanism
molecules
motion
n
neutralization
nine
opposite
orbital(s)
order
orientation
oxidation
oxidation number
oxidizing agent
p
Pauli exclusion
peak
photon(s)
precipitation
probability
product(s)
rate
rate law
reactant(s)
reaction mechanism
reaction rate
reducing agent
reduction
s
salt
Schrödinger
separate
solubility
soluble
solvation
speed
sublevels
titration
valence
water
wavelike

Aqueous Reactions *(5 marks)*

1. A substance is considered _____________ if it will dissolve in a specific solvent.
2. An _____________ in the oxidation number of an atom signifies oxidation, while a _____________ in the oxidation number signifies reduction.
3. A _____________ reaction is one in which the aqueous (dissolved) ions combine to form a solid that is not readily soluble.
4. _____________ is the addition of a known amount of solution to determine the volume (or concentration) of another solution.

Atomic Structure *(10 marks)*

5. The _____________ orbitals are only present in the fourth energy level and up.
6. Energy is _____________ when an electron moves from ground state to excited state.
7. Electron orbitals are areas of high _____________ of finding electrons.
8. Pauli said that two electrons can only occupy the same orbital if they have _____________ spins.
9. The electrons on the highest numbered subshells are the _____________ electrons.
10. The _____________ _____________ states that electrons must be added to the lowest energy state before adding electrons to a higher energy level.
11. The _____________ of a wave is the height of the wave from the origin to the top of the crest.
12. Einstein proposed that light is composed of particles, or quanta, known as _____________.
13. The quantum mechanical model of the atom was proposed by _____________.

Kinetics *(8 marks)*

14. The _____________ of a chemical reaction can be determined from the balanced chemical equation.
15. A reaction mechanism describes the _____________ in which bonds break and atoms rearrange in a chemical reaction.
16. If changing the concentration of a product does not change the rate, it does not appear in the _____________ _____________.
17. In an _____________ _____________, the particles are oriented in a favourable position that allows bonds to break and atoms to rearrange.
18. The difference between the energy at the peak of the activation diagram and the energy of the reactants is called the _____________ _____________.

Name: ____

Part C: Short Answer *(50 marks total)*

Answer each of the following questions 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 as well as the appropriate unit.

Aqueous Reactions *(25 marks)*

1. Give the oxidation number for each atom in Na_3AlF_6. Show your work. *(4 marks)*

   Na = +1; Al = +3; F = –1 *(1 mark each)*

   Show your work. *(1 mark)*

2. Complete the dissociation equation for the following ionic compound. *(2 marks)*

   $NaF_{(s)} \rightarrow \mathbf{Na^+_{(aq)} + F^-_{(aq)}}$

3. Balance the following redox reaction using the oxidation-number change method. You will earn up to 5 marks for the following steps:
   - 1 mark: assigning correct oxidation numbers
   - 1 mark: showing electron loss
   - 1 mark: showing electron gain
   - 1 mark: assigning correct coefficients
   - 1 mark: simplifying the terms that appear on both sides of the equation

   $Al_{(s)} + MnO_{2(s)} \rightarrow Al_2O_{3(s)} + Mn_{(s)}$

   $\mathbf{4Al_{(s)} + 3MnO_{2(s)} \rightarrow 2Al_2O_{3(s)} + 3Mn_{(s)}}$

4. Why must oxidation and reduction occur together? *(2 marks)*

   Oxidation and reduction must occur together because the number of electrons lost must equal the number of electrons gained. This assures that charge is conserved on both sides of the equation; therefore, an oxidation-reduction reaction, or redox reaction, is a reaction in which electrons are transferred (lost or gained) between reactants.

Name: ____

5. If 40.8 mL of 0.106 mol/L sulphuric acid neutralizes 61.8 mL of potassium hydroxide solution, answer the following questions to determine the concentration of the base
   a. Write the balanced chemical equation for the neutralization reaction. *(2 marks)*

$$\mathbf{H_2SO_{4(aq)} + 2\ KOH_{(aq)} \rightarrow K_2SO_{4(aq)} + 2\ H_2O_{(l)}}$$

   b. Use the concentration and volume of the known acid or base to calculate the moles of the known substance (acid or base). *(2 marks)*

$$\mathbf{\frac{0.106\ mol\ H_2SO_4}{L} \times 0.0408\ L = 0.00432\ mol\ H_2SO_4}$$

   c. Use the coefficients from the balanced equation to determine the moles of the unknown substance (acid or base). *(2 marks)*

$$\mathbf{\frac{mol\ H_2SO_4}{coefficient\ H_2SO_4} = \frac{mol\ KOH}{coefficient\ KOH}}$$

$$\mathbf{\frac{0.00432\ mol\ H_2SO_4}{1\ mol\ H_2SO_4} = \frac{mol\ KOH}{2\ mol\ KOH}}$$

$$\mathbf{0.00864\ mol\ KOH = mol\ KOH}$$

d. Calculate the required concentration of the base. *(2 marks)*

$$\mathbf{C = \frac{moles}{volume} = \frac{0.00864\ mol\ KOH}{0.0618\ L} = 0.140\ mol/L}$$

6. Write the balanced neutralization reaction given the following reactants. *(2 marks)*

   $2\ HCl + Mg(OH)_2 \rightarrow \mathbf{2\ H_2O + MgCl_2}$

7. Use the solubility rules to predict whether a precipitate will form when the following pair of aqueous solutions are mixed. If a precipitate forms, write the balanced net ionic equation for the reaction. *(2 marks)*

   $Pb(C_2H_3O_2)_{2(aq)}$ and $LiCl_{(aq)}$

   Precipitate will form.
   $\mathbf{Pb^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)} \rightarrow PbCl_{2(s)}}$

Name: ____

Atomic Structure ***(8 marks)***

8. Write the electron configuration and noble gas configuration for a chlorine atom. *(2 marks)*

   electron configuration: $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^5}$ *(1 mark)*

   noble gas configuration: $\mathbf{[Ne]\ 3s^2 3p^5}$ *(1 mark)*

9. Outline the contributions of Neils Bohr, Werner Heisenberg, and Erwin Schrödinger in relation to our understanding of the quantum mechanical model of the atom. *(3 marks)*

   Neils Bohr interpreted the Bohr model of the atom, whereby he proposed that electrons are found in specific circular orbits around the nulceus of an atom. *(1 mark)*

   This led to the development of the Heisenberg uncertainty principle by Werner Heisenberg, which states that it is impossible to know the exact location of an electron and where it is going at the same time. *(1 mark)*

   Erwin Schrödinger used complex math to develop an equation that represents the statistical probability of finding an electron in a particular space in the atom. *(1 mark)*

10. Write the electronic configurations for the valence electrons of each of the following. *(3 marks)*

    a. Mg^{2+}

    $2s^22p^6$

    b. S^{2-}

    $3s^23p^6$

    c. sodium

    $3s^1$

Kinetics ***(17 marks)***

11. For the reaction $CH_3COCH_3 + I_2 \rightarrow CH_3COCH_2I + HI$, the following data was obtained:

Experiment	Initial $[CH_3COCH_3]$	Initial $[I_2]$	Reaction Rate (mol/L · s)
1	0.100 mol/L	0.100 mol/L	1.16×10^{-7}
2	0.0500 mol/L	0.100 mol/L	5.18×10^{-8}
3	0.0500 mol/L	0.500 mol/L	5.18×10^{-8}

    a. Write the rate expression for the reaction. *(1 mark)*

    Rate = $K[CH_3COCH_3]^1[I_2]^0$

    b. What order is the reaction with respect to each of the reactants? *(2 marks)*

    With respect to CH_3COCH_3, the order is 1.
    With respect to I_2, the order is 0.

Name: ____

c. What is the overall order of this reaction? *(1 mark)*

Overall order is 1 + 0 = 1 (1st order)

d. Calculate the specific rate constant. *(2 marks)*

$1.16 \times 10^{-7} = k[0.100]^1$

$k = 1.16 \times 10^{-6}$

e. What is the reaction rate if the initial concentrations are both 0.045 mol/L? *(2 marks)*

Rate = $k[0.045]^1 = (1.16 \times 10^{-6})(0.045) = 5.22 \times 10^{-8}$

12. List two ways in which reaction rates can generally be increased. *(2 marks)*

Any two of the following can be listed for 1 mark each: increase the reactant concentration, decrease the volume of the vessel containing gaseous reactants, increase the temperature, decrease the particle size, and add a catalyst.

13. Consider the following reaction mechanism.

$H_2O_2 + I^- \rightarrow H_2O + IO^-$ (slow) $\Delta H = 45$ kJ

$H_2O_2 + IO^- \rightarrow H_2O + O_2 + I^-$ (fast) $\Delta H = 35$ kJ

a. What is (are) the intermediate(s)? *(1 mark)*

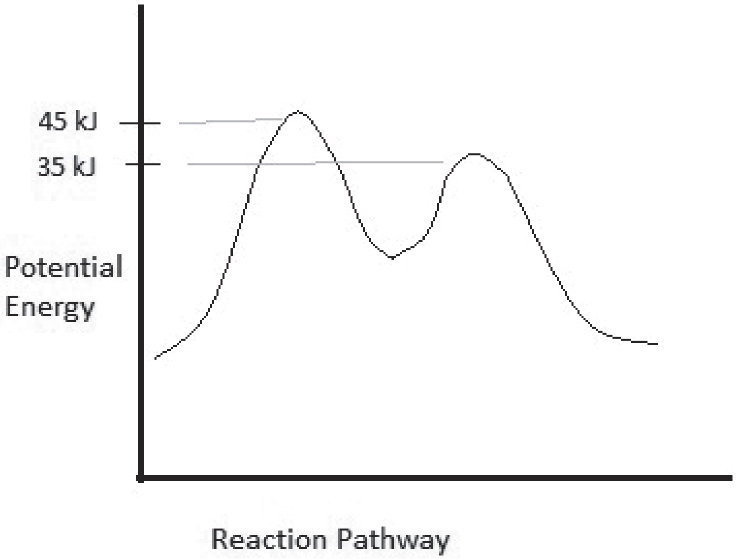
IO^-

b. What is the overall equation? *(2 marks)*

$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

c. Draw a potential energy diagram for the reaction. *(3 marks)*

0.5 mark for the "PE" label on the *y*-axis
0.5 mark for the "Reaction Pathway" label on the *x*-axis
1 mark for the activated complex at 45 kJ
1 mark showing the activated complex of fast step to be 35 kJ

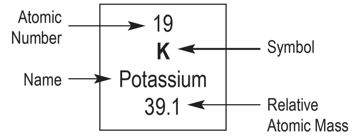


d. In which step would a catalyst be most useful? *(1 mark)*

It would be most useful in step 1 (the slowest step).

Name: ____

Periodic Table of Elements



	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)

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)	Irdium	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)	Sulphur	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
Einstienium	(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		

Name: ____________________________________

Table of Electronegativities

	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** —	1
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** —	2
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** —	3
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** —	4
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** —	5
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** —	6
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** —	7

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** —

Relative Strengths of Acids Table

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulphuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4×10^{-2}
Sulphurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7×10^{-2}
Hydrogen sulphate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3×10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1×10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0×10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3×10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7×10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1×10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7×10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5×10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6×10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4×10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4×10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4×10^{-7}
Hydrogen sulphide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3×10^{-8}
Hydrogen sulphite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2×10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7×10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7×10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0×10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4×10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4×10^{-13}
Hydrogen sulphide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2×10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8×10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Name: ____

Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	low solubility
	all others	soluble
sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulphide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulphite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

Table of Standard Reduction Potentials with Values

Oxidized Species	↔	Reduced Species	E°/V
$Li^+_{(aq)} + e^-$	↔	$Li_{(s)}$	–3.04
$K^+_{(aq)} + e^-$	↔	$K_{(s)}$	–2.93
$Ca^{2+}_{(aq)} + 2e^-$	↔	$Ca_{(s)}$	–2.87
$Na^+_{(aq)} + e^-$	↔	$Na_{(s)}$	–2.71
$Mg^{2+}_{(aq)} + 2e^-$	↔	$Mg_{(s)}$	–2.37
$Al^{3+}_{(aq)} + 3e^-$	↔	$Al_{(s)}$	–1.66
$Mn^{2+}_{(aq)} + 2e^-$	↔	$Mn_{(s)}$	–1.19
$H_2O_{(l)} + e^-$	↔	$\frac{1}{2}H_{2(g)} + OH^-_{(aq)}$	–0.83
$Zn^{2+}_{(aq)} + 2e^-$	↔	$Zn_{(s)}$	–0.76
$Fe^{2+}_{(aq)} + 2e^-$	↔	$Fe_{(s)}$	–0.45
$Ni^{2+}_{(aq)} + 2e^-$	↔	$Ni_{(s)}$	–0.26
$Sn^{2+}_{(aq)} + 2e^-$	↔	$Sn_{(s)}$	–0.14
$Pb^{2+}_{(aq)} + 2e^-$	↔	$Pb_{(s)}$	–0.13
$H^+_{(aq)} + e^-$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}_{(aq)} + e^-$	↔	$Cu^+_{(aq)}$	+0.15
$SO_4^{2-}{}_{(aq)} + 4\ H^+_{(aq)} + 2e^-$	↔	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}_{(aq)} + 2e^-$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^-$		$2OH^-_{(aq)}$	+0.40
$Cu^+_{(aq)} + e^-$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^-$	↔	$I^-_{(aq)}$	+0.54
$Fe^{3+}_{(aq)} + e^-$	↔	$Fe^{2+}_{(aq)}$	+0.77
$Ag^+_{(aq)} + e^-$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^-$	↔	$Br^-_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^-$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e^-$	↔	$2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^-$	↔	$Cl^-_{(aq)}$	+1.36
$MnO_4^-{}_{(aq)} + 8H^+_{(aq)} + 5e^-$	↔	$Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^-$	↔	$F^-_{(aq)}$	+2.87

Name: ____

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

Names, Formulas, and Charges of Common Ions *(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^-	sulphate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulphate	HSO_4^-
fluoride	F^-	sulphide	S^{2-}
hydroxide	OH^-	hydrogen sulphide	HS^-
hypochlorite	ClO^-	sulphite	SO_3^{2-}
iodide	I^-	hydrogen sulphite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Name: ____

Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
$NH4^+$	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	$Pb2^{2+}$	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

Common Ions *(continued)*

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- $(C_2H_3O_2^-)$	Acetate (or ethanoate)	HS^-	Hydrogen sulphide	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-}	Sulphate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulphide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulphite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulphate
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 sulphate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Name: ____

Midterm Practise Examination Answer Sheet

Part A: Multiple Choice *(27 marks)*

For each multiple-choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Example: Ⓐ Ⓑ ● Ⓓ

1. Ⓐ Ⓑ ● Ⓓ
2. Ⓐ ● Ⓒ Ⓓ
3. Ⓐ Ⓑ Ⓒ ●
4. Ⓐ Ⓑ Ⓒ ●
5. ● Ⓑ Ⓒ Ⓓ
6. Ⓐ ● Ⓒ Ⓓ
7. Ⓐ Ⓑ Ⓒ ●
8. Ⓐ ● Ⓒ Ⓓ
9. ● Ⓑ Ⓒ Ⓓ
10. Ⓐ ● Ⓒ Ⓓ
11. Ⓐ Ⓑ Ⓒ ●
12. Ⓐ Ⓑ Ⓒ ●
13. Ⓐ ● Ⓒ Ⓓ
14. Ⓐ Ⓑ ● Ⓓ
15. ● Ⓑ Ⓒ Ⓓ
16. Ⓐ Ⓑ ● Ⓓ
17. ● Ⓑ Ⓒ Ⓓ
18. Ⓐ Ⓑ ● Ⓓ
19. Ⓐ Ⓑ ● Ⓓ
20. ● Ⓑ Ⓒ Ⓓ
21. Ⓐ ● Ⓒ Ⓓ
22. Ⓐ Ⓑ ● Ⓓ
23. Ⓐ ● Ⓒ Ⓓ
24. Ⓐ ● Ⓒ Ⓓ
25. ● Ⓑ Ⓒ Ⓓ
26. Ⓐ Ⓑ ● Ⓓ
27. Ⓐ Ⓑ ● Ⓓ

Part B: Fill-in-the-Blank *(23 marks)*

For each fill-in-the-blank question, write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Aqueous Reactions (*5 marks*)

1. soluble
2. increase

   decrease
3. precipitation
4. titration

continued

Atomic Structure ***(10 marks)***

5. f
6. absorbed
7. probability
8. opposite
9. valence
10. Aufbau
principle
11. amplitude
12. photons
13. Schrödinger

Kinetics ***(8 marks)***

14. rate
15. order
16. rate
law
17. effective
collision
18. activation
energy

GRADE 12 CHEMISTRY (40S)

Module 4
Chemical Equilibrium

MODULE 4: CHEMICAL EQUILIBRIUM

Introduction to Module 4

So far, you have assumed that reactions proceed from reactants to products as indicated by balanced molecular equations. Very few reactions actually proceed to completion. Many reactions actually proceed in both directions, left to right **and** right to left. At equilibrium, the concentration of both reactants and products remains constant because the forward and reverse reactions are occurring at the same rate.

In this module, you will examine why and when equilibrium exists, as well as the concentrations of products and reactants at equilibrium. You will also determine whether or not a system is at equilibrium and, if not, predict what the system will do to attain equilibrium.

Note

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.

Assignments in Module 4

Note

Please remember to work on Assignment 1.9: Researching a Redox Application as you work through this course. Please submit it at the end of Module 6.

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	Dynamic Equilibrium
2	Assignment 4.2	Writing Equilibrium Law
3	Assignment 4.3	Solving K_{eq} Problems
4	Assignment 4.4	Le Châtelier's Principle
5	Assignment 4.5	Interpreting a Concentration-versus-Time Graph
6	Assignment 4.6	The Haber Process
7	Assignment 4.7	Working with Solubility Product Constant
8	Assignment 4.8	Kidney Stones

LESSON 1: WHAT IS EQUILIBRIUM? (1 HOUR)

Lesson Focus

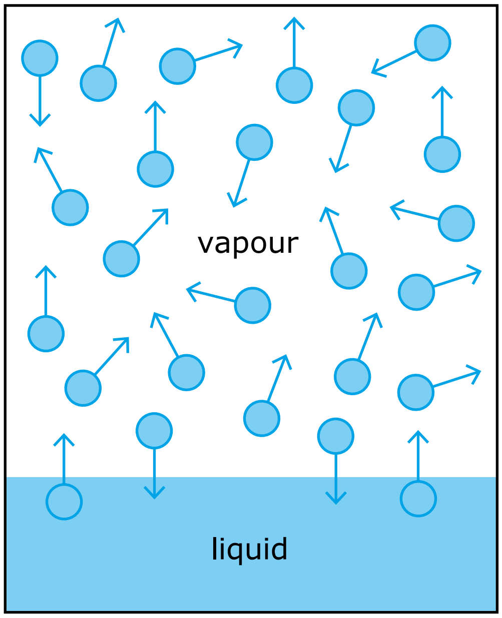
SLO C12-4-01: Relate the concept of equilibrium to physical and chemical systems.
Include: conditions necessary to achieve equilibrium

Lesson Introduction

The amount of water on Earth (including the atmosphere) stays relatively constant. Water evaporates or ice sublimates into the gas phase and later returns to Earth in the form of precipitation and dew. This balance between the various phases of water is known as the water cycle, and is an example of equilibrium.

Do you play any team sports? During the game, players enter and leave the field, rink, or court; however, the number of players on the playing surface never changes. This is another example of a system in equilibrium. In this lesson, you will study types of equilibrium systems in the chemical world.

Understanding the equilibrium for a reaction allows chemists to manipulate the conditions in order to control the speed, direction, and efficiency of the process. Alzheimer's disease has been linked to the equilibrium of specific proteins. Research is being done on how to manipulate that equilibrium to prevent or stop the spread of the disease.



Dynamic Equilibrium

In Grade 11 Chemistry, you learned about equilibrium in terms of physical changes.

For example, when liquid in an open container is left to sit, it will eventually evaporate to dryness. This is due to particles on the surface of the liquid vapourizing and leaving the container. However, if the container is sealed, the vapour is trapped and cannot leave the container. Eventually, the space above the liquid becomes saturated with vapour, and for every molecule that evaporates, another condenses. You might recall studying an illustration similar to the one shown in Grade 11 Chemistry.

Liquid in a closed container at a constant temperature reaches equilibrium with its vapour. At equilibrium,

> the rate of evaporation is equal to the rate of condensation

The pressure created by the vapour at equilibrium is known as the **vapour pressure**, abbreviated P_{vap}. This is not the only example of physical equilibrium. Another example is a saturated solution.

This liquid-vapour system is more accurately an example of **dynamic equilibrium**. Even though the amount of vapour and liquid do not change, there is constant movement between the two states. The amounts remain constant because the rate of the forward process (evaporation) is equal to the rate of the reverse process (condensation). As one liquid molecule vapourizes, another vapour molecule condenses simultaneously. How can you show this dynamic equilibrium? Using a double arrow ($\rightleftharpoons$) indicates the reaction is reversible; that is, it can proceed both left to right and right to left. In a system at equilibrium, the reaction proceeds in **both** directions simultaneously. Chemists describe the liquid-vapour equilibrium as follows:

$$H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$$

Conditions for Equilibrium

Dynamic equilibrium cannot be established in an open container because the vapour molecules can escape. Imagine a scenario where hockey players went home after leaving the ice following a shift change. Eventually, you would run out of players to play the opposing team. It would certainly not be a two-way reaction!

Equilibrium can only be established when all particles are kept in a sealed container and certain conditions are kept constant. This way, the vapour particles do not escape; instead, they return to the liquid state. A system where no particles are allowed to enter or leave and conditions are kept constant is known as a **closed system**. Equilibrium can only occur in a closed system when pressure and temperature are kept constant.

Investigate another liquid-vapour equilibrium online by using a video hosting site such as www.youtube.com/ and searching for an equilibrium that is taking place in a closed system with a constant temperature.

After watching such an animation, you should notice that each time a liquid molecule enters the gas phase, a gaseous molecule enters the liquid phase.

The system is in dynamic equilibrium. If the temperature is changed, the relative number of particles in each phase will become different and may reach a new equilibrium. To maintain equilibrium, it is important to maintain a constant temperature.

Characteristics of Chemical Equilibrium

So far, your study of chemical reactions has focused on reactions that proceed completely from reactants to products; however, many chemical reactions are reversible. Consider the combination of nitrogen gas and oxygen gas to form nitrogen monoxide gas.

$$N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$$

Watch an animation of the above reaction online at www.dlt.ncssm.edu/tiger/Flash/equilibrium/N2-O2_Equilibrium.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm
2. scrolling down the page until you find the animation titled "**N2-O2_Equilibrium.html**" or "**N2-O2_Equilibrium.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Nitrogen monoxide is emitted in abundance from high temperature combustion in the air. This gas is a catalyst in the reactions that cause the destruction of ozone. In the animation (link provided above), nitrogen gas and oxygen gas are introduced into a sealed container. The oxygen gas (red) combines with the nitrogen gas (blue) to form two molecules of nitrogen monoxide. The equilibrium state is characterized by constant concentrations of both reactants (N_2 and O_2) and products (NO). As such, the reaction does **not** go completely to products, but rests somewhere in between.

Learning Activity 4.1: Equilibrium Terms

1. Match the terms with the correct description.

___	It is the formation of products from reactants.	a. reversible reaction
___	It is a chemical reaction in which the products can return to the original reactants.	b. chemical equilibrium
___	It is the speed at which a reaction occurs.	c. reverse reaction
___	It is the formation of reactants from products.	d. reaction rate
___	It occurs when the rate of formation of products is equal to the rate of formation of reactants over time.	e. forward reaction

2. Does a reaction stop once it has reached equilibrium? Explain.
3. Why is chemical equilibrium referred to as dynamic equilibrium?

Check the answer key.

Lesson Summary

In this lesson, you learned about the concept of equilibrium. You now know what conditions are necessary to achieve equilibrium. The following is a short summary of the lesson.

- The double arrow indicates the reaction is reversible.
- In reversible reactions,
  - conversion to products is the forward reaction
  - conversion to reactants is the reverse reaction
- For a system at equilibrium, the rate of the forward reaction equals the rate of the reverse reaction.
- The equilibrium state is dynamic, since **both** forward and reverse reactions occur simultaneously.

In the next lesson, you will learn about a more mathematical approach to equilibrium. This approach involves writing equilibrium law expressions from balanced chemical equations.

Notes

Assignment 4.1: Dynamic Equilibrium (5 marks)

1. In the reaction below, which is the forward reaction and which is the reverse reaction? *(2 marks)*

   $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$

2. Why will a chemical reaction only reach equilibrium in a closed container? *(1 mark)*

3. When considering most chemical reactions, it is assumed that reactants are entirely consumed. Is this assumption accurate? *(2 marks)*

Notes

LESSON 2: EQUILIBRIUM LAW EXPRESSIONS (2 HOURS)

Lesson Focus

SLO C12-4-02: Write equilibrium law expressions from balanced chemical equations for heterogeneous and homogeneous systems.
Include: mass action expression

Lesson Introduction

Now that you know something about equilibrium, you will go one step further and learn to use the **mass action expression**. This mathematical approach to equilibrium shows the relationship between the concentration of reactants and products at equilibrium. Using this, you can calculate the **equilibrium constant**.

Equilibrium Law

In 1864, Cato Guldberg and Peter Waage proposed the **law of mass action**, also known as the **equilibrium law**. They studied many systems at equilibrium, and discovered a relationship between the concentration of reactants and products at equilibrium. They proposed that equilibrium law be a ratio of product concentrations to reactant concentrations. The value of this ratio is called the **equilibrium constant**.

For the reaction,

$$aA + bB \rightleftharpoons cC + dD$$

A, B, C, and D are reactants and products, and the lower case letters *a, b, c,* and *d* are the molar coefficients in a balanced reaction. If the forward and reverse reactions were elementary reactions, the rate laws for both forward and reverse reactions would be

$$\text{rate}_{\text{forward}} = K_{\text{f}}[\text{C}]^{c}[\text{D}]^{d}$$

and

$$\text{rate}_{\text{reverse}} = K_{\text{r}}[\text{A}]^{a}[\text{B}]^{b}$$

Since $\text{rate}_{\text{forward}} = \text{rate}_{\text{reverse}}$ at equilibrium, we can state that

$$K_r[A]^a[B]^b = K_f[C]^c[D]^d$$

By rearranging the expression, you can solve for the ratio of rate constants as follows:

$$\frac{K_f}{K_r} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The ratio of rate constants is condensed to one constant, K_{eq}, called the equilibrium constant. The **law of mass action**, or equilibrium law, is the ratio of the concentration of products to reactants at equilibrium. The value of this ratio is the equilibrium constant. If we replace the ratio of the rate constants with K_{eq}.

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

$$\text{or } K_{eq} = \frac{[\text{products}]}{[\text{reactants}]}$$

Note

K_{eq} is often written as K_c or just K.

Writing the Equilibrium Law

Homogeneous equilibrium is achieved when all reactants and products are in the same phase: gaseous $_{(g)}$ or aqueous $_{(aq)}$.

Example 1

Write the equilibrium law for the following equation:

$$N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$$

Step 1: Write the product concentrations over the reactant concentrations.

Step 2: Convert each coefficient from the balanced chemical equation to an exponent. For example, the coefficient for N_2 is 1, so it is written as $[N_2]$ in the mass action expression. The coefficient for H_2 is 3, so it is written as $[H_2]^3$ and NH_3 is written as $[NH_3]^2$.

Since all reactants and products are gaseous, the equilibrium law would be

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Heterogeneous equilibrium refers to reactions involving reactants and products in more than one state. When writing the mass action expression or equilibrium law, substances that are solids $_{(s)}$ or liquids $_{(l)}$ are omitted. This is because concentrations of solids and liquids are constant and are, therefore, ignored.

Example 2

Write the equilibrium law for the following equation:

$$C_{(s)} + H_2O_{(g)} \rightleftharpoons CO_{(g)} + H_{2(g)}$$

Step 1: Write the product concentrations over the reactant concentrations.

Step 2: Convert each coefficient from the balanced chemical equation to an exponent. The coefficient for H_2O is 1, so it is written as $[H_2O]$ in the mass action expression. The coefficient for H_2 is 1, so it is written as $[H_2]$ and CO is written as $[CO]^1$, because the coefficient for CO is 1. Since C is in the solid state, it is omitted from the equilibrium law.

$$K_{eq} = \frac{[H_2][CO]}{[H_2O]}$$

The Reaction Quotient

When reactants and products are both present in a mixture, it is useful to determine if the system is at equilibrium or not. The reaction quotient, Q, is calculated to determine whether the reaction is at equilibrium or not. The reaction quotient can also identify in which direction a reaction will proceed if the system is not at equilibrium. Calculating the reaction quotient will follow in Lesson 3.

Learning Activity 4.2: The Law of Chemical Equilibrium

1. Match the terms with the correct description below. In some cases, there is more than one correct answer.

___	It is the equilibrium condition for a chemical reaction involving substances in more than one state.	a. equilibrium position
___	It is used to determine if a reaction has reached equilibrium.	b. law of chemical equilibrium
___	It depends on the initial concentrations of the substances in a reaction.	c. reaction quotient
___	It states that every reaction proceeds to an equilibrium state with a specific K_{eq}.	d. homogeneous equilibrium
___	It expresses the relative concentration of reactants and products at equilibrium in terms of an equilibrium constant.	e. law of mass action
___	It is the equilibrium condition for reactions in which products and reactants are in the same state.	f. heterogeneous equilibrium
___	It is the ratio of product concentration to reactant concentration at equilibrium.	g. equilibrium constant

2. Write the equilibrium law (mass action expression) for each of the following reactions.
   a. $SO_{2(g)} + NO_{2(g)} \rightleftharpoons SO_{3(g)} + NO_{(g)}$
   b. $2\,C_{(s)} + 3\,H_{2(g)} \rightleftharpoons C_2H_{6(g)}$
   c. $3\,O_{2(g)} \rightleftharpoons 2\,O_{3(g)}$
   d. $MgCO_{3(s)} \rightleftharpoons CO_{2(g)} + 2\,MgO_{(s)}$
   e. $NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$

Check the answer key.

Lesson Summary

In this lesson, you learned how to write equilibrium law expressions from balanced chemical equations for both heterogeneous and homogeneous systems. Now that you know how to set up mass action expressions, in the next lesson, you will solve problems involving equilibrium constants.

Notes

Assignment 4.2: Writing Equilibrium Law (10 marks)

1. Write the equilibrium law (mass action expression) for each of the following reactions. *(2 marks each)*

   a. $I_{2(aq)} \rightleftharpoons I_{2(s)}$

   b. $Cl_{2(g)} + PCl_{3(g)} \rightleftharpoons PCl_{5(g)}$

   c. $I_{2(g)} + Cl_{2(g)} \rightleftharpoons 2\,ICl_{(g)}$

   d. $2\,NO_{2(g)} \rightleftharpoons 2\,NO_{(g)} + O_{2(g)}$

   e. $2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)}$

Notes

LESSON 3: THE EQUILIBRIUM CONSTANT (3 HOURS)

Lesson Focus

SLO C12-4-03: Use the value of the equilibrium constant (K_{eq}) to explain how far a system at equilibrium has gone towards completion.

SLO C12-4-04: Solve problems involving equilibrium constants.

Lesson Introduction

In the previous lesson, you learned how to set up mass action expressions. You were also introduced to the notion of the equilibrium constant. In this lesson, you will put that knowledge to use by interpreting the value of the equilibrium constant and by solving problems involving K_{eq}.

Equilibrium Constant

The **equilibrium constant** is the ratio of product concentrations to reactant concentrations at equilibrium. You may find this constant abbreviated as K_{eq}, K_c, or just K. The only factor affecting K is **temperature**. Equilibrium constants are specific for only **one reaction** at a particular temperature.

The equilibrium constant can indicate whether products or reactants are favoured at equilibrium. In other words, based on the value of the K_{eq} at equilibrium, you can determine whether there are more products or reactants at equilibrium. The following are the three possible scenarios:

1. If $\boldsymbol{K = 1}$, the ratio of [products] = [reactants] at equilibrium.

   Neither reactants nor products are favoured.

2. If $\boldsymbol{K > 1}$, the value of product concentrations raised to the power of their coefficients is greater than reactant concentration at equilibrium.

   In other words, the **numerator** of the equilibrium law is greater than the **denominator**. Since product concentrations are in the numerator, this reaction is in equilibrium when there is more product. We state that *products are favoured* and the *equilibrium position lies to the right*. We also say the *forward reaction is favoured*.

For example, if the reaction $A + B \rightleftharpoons C + D$ has a $K_{eq} = 1 \times 10^5$, mixing A and B results in almost a complete conversion to products C and D.

3. If $K < 1$, the value of product concentrations is less than reactant concentrations at equilibrium.

   When $K_{eq} < 1$, the numerator of the equilibrium law is less than the denominator. Since the reactants are in the denominator, this reaction is at equilibrium when there are more reactants present. We say that reactants are favoured; the equilibrium position lies to the left and the reverse reaction is favoured.

   For example, if the reaction $E + F \rightleftharpoons G + H$ has a $K_{eq} = 1 \times 10^{-5}$, mixing E and F results in almost no conversion to products.

Finding K_{eq}

You can calculate the value of K_{eq} if you are given the concentrations of reactants and products at equilibrium. The values are then substituted into the equilibrium law.

Example 1

For the reaction

$$N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$$

at 225°C, a 2.0 L container holds 0.040 moles of N_2, 0.15 moles of H_2 and 0.50 moles of NH_3. If the system is at equilibrium, calculate K_{eq}.

Step 1: Change all quantities into concentrations (mol/L).

$$\text{Concentration} = \frac{\text{moles}}{\text{litres}}$$

$$C_{N_2} = \frac{0.040 \text{ moles}}{2.0 \text{ L}} = 0.020 \text{ mol/L}$$

$$C_{H_2} = \frac{0.15 \text{ moles}}{2.0 \text{ L}} = 0.075 \text{ mol/L}$$

$$C_{NH_3} = \frac{0.50 \text{ moles}}{2.0 \text{ L}} = 0.25 \text{ mol/L}$$

Step 2: Write the equilibrium law for the reaction. Remember that this expression only includes products and reactants in the aqueous and gaseous phases.

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Step 3: Substitute the concentration values and calculate K_{eq}. Do not forget to include the exponents in your calculation.

$$K_{eq} = \frac{(0.25)^2}{(0.020)(0.075)^3}$$

$$K_{eq} = \frac{0.0625}{(0.020)(4.219 \text{ x } 10^{-4})}$$

$$K_{eq} = 7.4 \text{ x } 10^3$$

In the above example, the forward reaction is favoured, as the value of K_{eq} is greater than one. Another quick check that can be performed to verify that the value of K_{eq} is correct is to look at the concentration of the reactants and the products at equilibrium. Since K_{eq} is larger than one, it makes sense that the concentration of the products is greater than the concentration of the reactants. Finally, there are no units for K_{eq}.

Finding Concentrations

Another type of problem involving equilibrium constants may require you to solve for an equilibrium concentration. In such cases, K_{eq} and the equilibrium concentrations of all remaining reactants and products must be given.

Example 2

For the following reaction at 210°C, the K_{eq} is 64.0.

$$N_{2(g)} + O_{2(g)} \rightleftharpoons 2\,NO_{(g)}$$

The equilibrium concentrations of N_2 and O_2 are 0.40 mol/L and 0.60 mol/L, respectively. Calculate the equilibrium concentration of NO.

Step 1: Write out the equilibrium law.

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$

Step 2: Using algebra, rearrange the equilibrium law to solve for [NO].

$$[\mathrm{NO}]^2 = K_{\mathrm{eq}}[\mathrm{N_2}][\mathrm{O_2}]$$
$$[\mathrm{NO}] = \sqrt{K_{\mathrm{eq}}[\mathrm{N_2}][\mathrm{O_2}]}$$

Step 3: Substitute the given K_{eq} and equilibrium concentrations and then solve for the unknown. Since you are solving for a concentration, your answer requires the units of mol/L.

$$[\mathrm{NO}] = \sqrt{(64.0)(0.40)(0.60)} = \sqrt{15.36}$$
$$[\mathrm{NO}] = 3.9\ \mathrm{mol/L}$$

At equilibrium, the concentration of NO is 3.9 mol/L.

The ICE Table

Another problem type involves putting known amounts of reactants together and calculating the concentrations after the reaction has reached equilibrium. An ICE table is a convenient way to organize data for equilibrium problems. The letter *I* stands for the **initial** concentrations of the reactants and products. The letter *C* stands for the **change** in reactants and products from the initial conditions to equilibrium. The letter *E* represents the concentrations of the reactants and products at **equilibrium**. Remember that concentration changes in solids and liquids are very slight and are, for that reason, ignored.

To set up the table, write the balanced chemical equation widely spaced and create columns under each chemical. Write the letters *I*, *C*, and *E* on the left-hand side of the table. Insert the known values (the initial concentrations that you are given) in the table and use x to represent the unknown value.

Here are some things you should notice about the ICE table.

1. The initial concentrations of the reactants placed in the vessel are greater than their equilibrium concentrations because these species have been consumed to make the product.
2. The products should have higher concentrations at equilibrium because, initially, they were not present in the vessel.
3. All aqueous and gaseous species at equilibrium should have non-zero concentrations.

It is always a good idea to review the final numbers in an ICE table to ensure that the values make sense. For example, it would not make sense for the concentration of a reactant gas to be higher at equilibrium than it was initially.

Example 3

0.500 mol H_2 and 0.500 mol I_2 are placed in a 1.00 L flask at 430°C. The equilibrium constant, K_{eq}, is 54.3 at this temperature. Calculate the concentrations of H_2, I_2, and HI at equilibrium.

Step 1: Set up the ICE table and fill in known values.

Since both reactants are used, the change in concentration is indicated as $-x$. It is not known exactly how much reactant is used to make the product (we will assign this value as x); however, you can assume that both reactants are used in the same proportions because the balanced reaction (below) shows a 1:1 relationship between products. The negative value indicates that the reactant is used, as opposed to being formed.

Since 2 moles of product are formed, the concentration change is $2x$. The positive value indicates that product is formed, as opposed to being used.

	$H_{2(g)}$	$+ I_{2(g)}$ $\rightarrow$	$2HI_{(g)}$
I	0.500 mol/L	0.500 mol/L	0 mol/L
C	$-x$	$-x$	$2x$
E	0.500 mol/L $-x$	0.500 mol/L $-x$	$2x$

Step 2: Write the equilibrium expression for the reaction.

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

Step 3: Substitute the values at equilibrium into the equilibrium expression.

$$54.3 = \frac{(2x)^2}{(0.500 - x)(0.500 - x)} = \frac{(2x)^2}{(0.500 - x)^2}$$

Next, take the square root of both sides.

$$\sqrt{54.3} = \sqrt{\frac{(2x)^2}{(0.500 - x)^2}}$$

$$7.37 = \frac{2x}{0.500 - x}$$

Step 4: Rearrange the equation and solve for x.

$$7.37\,(0.500 - x) = 2x$$

$$3.685 - 7.37x = 2x$$

$$3.685 = 2x + 7.37x$$

$$3.685 = 9.37x$$

$$x = \frac{3.685}{9.37} = 0.393\ \text{mol/L}$$

At equilibrium, the concentration of H_2, I_2, and HI are (using the E line in the ICE table)

$$[H_2] = 0.500\ \text{mol/L} - x = 0.500\ \text{mol/L} - 0.393\ \text{mol/L} = 0.107\ \text{mol/L}$$

$$[I_2] = 0.500\ \text{mol/L} - x = 0.500\ \text{mol/L} - 0.393\ \text{mol/L} = 0.107\ \text{mol/L}$$

$$[HI] = 2\ x = (2)0.393\ \text{mol/L} = 0.786\ \text{mol/L}$$

There are slightly different variations of this problem-solving method. Here is another example using the ICE table in which the objective is to calculate the equilibrium constant given the initial concentrations.

Example 4

For the reaction

$$H_{2(g)} + F_{2(g)} \rightleftharpoons 2HF_{(g)}$$

1.00 mole of hydrogen and 1.00 mole of fluorine are sealed in a 1.00 L flask at 150°C and allowed to react. At equilibrium, 1.32 moles of HF are present. Calculate the equilibrium constant using the ICE method.

Step 1: Set up the ICE table and fill in known values.

The flask is 1.00 L, $[H_2] = [F_2] = 1.00$ mol/L, and the initial concentration of HF is zero (because it is a product).

	$H_{2(g)}$	$+\ F_{2(g)}$	$\rightarrow\ 2HF_{(g)}$
I	1.00	1.00	0.00
C			
E			

Step 2: Determine the concentration of H_2 and F_2 that will react. From the equation above, you can see that when 1 mole of hydrogen reacts with 1 mole of fluorine, 2 moles of product are formed.

From the question, we know that the hydrogen and fluorine will react to form an equilibrium concentration of 1.32 mol/L HF.

$$1.32\ \text{mol/L HF}\left(\frac{1\ \text{mole}\ H_2}{2\ \text{moles HF}}\right) = 0.66\ \text{mol/L}\ H_2$$

Therefore, 0.66 mol/L of H_2 and F_2 will react.

Step 3: Calculate the equilibrium concentrations. The data from the previous step is recorded as the change (C) by subtracting 0.66 mol/L from H_2 and F_2 and adding 1.32 mol/L HF.

	$H_{2(g)}$	$+\ F_{2(g)}$	$\rightarrow\ 2HF_{(g)}$
I	1.00	1.00	0.00
C	−0.66	−0.66	+1.32
E	0.34	0.34	1.32

This means the equilibrium concentrations of H_2 and F_2 are 0.34 mol/L and HF is 1.32 mol/L.

Step 4: Substitute the equilibrium concentrations into the equilibrium law and solve for K_{eq}.

$$K_{eq} = \frac{[HF]^2}{[H_2][F_2]} = \frac{(1.32)^2}{(0.34)(0.34)} = \frac{1.74}{0.1156} = 15.07$$

Reaction Quotient Calculations

To verify whether the reaction has reached equilibrium, the reaction quotient can be calculated and compared to a known K_{eq}. Consider the following reaction:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

The reaction quotient is calculated the same way as K_{eq}.

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

One of three possible outcomes can result once the value of Q is known.

1. **Q is equal to K_{eq}.** This means that the system is at equilibrium. The concentrations listed are the same as the concentrations at equilibrium.
2. **Q is greater than K_{eq}.** If Q is a larger number than K_{eq}, then the ratio of [products] to [reactants] is larger than the ratio should be at equilibrium. This means that the system is not at equilibrium and, in this case, the reaction must produce more reactant to attain a value of Q = K_{eq}. A net change of products to make more reactants is needed; therefore, the reverse reaction is favoured.
3. **Q is less than K_{eq}.** If Q is a smaller number than K_{eq}, then the ratio of [products] to [reactants] is smaller than the ratio should be at equilibrium. This means that the system is not at equilibrium and, in this case, the reaction must produce more product to attain a value of Q = K_{eq}. A net change of reactants to make more product is needed; therefore, the forward reaction is favoured.

Example 5:

Given the following reaction,

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

the system is analyzed at 500°C and reveals the following concentrations of reactants and products.

$[NH_3] = 1.0 \times 10^{-3}$ mol/L $[N_2] = 1.0 \times 10^{-5}$ mol/L $[H_2] = 2.0 \times 10^{-3}$ mol/L

The value of K_{eq} at this temperature is 6.0×10^{-2}

1. Calculate the reaction quotient Q.

   Set up the reaction quotient expression and fill in known values to calculate Q.

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{[1.0 \times 10^{-3}]^2}{[1.0 \times 10^{-5}][2.0 \times 10^{-3}]^3} = \frac{1.0 \times 10^{-6}}{8.0 \times 10^{-14}} = 1.25 \times 10^{7}$$

   Use the value of Q to determine whether the system is at equilibrium or not.

   Since the value of Q = 1.25×10^7 is much larger than $K_{eq} = 6.0 \times 10^{-2}$, the system is not at equilibrium. To be at equilibrium, Q = K_{eq}.

2. If it is not at equilibrium, determine which way the reaction must proceed to attain equilibrium for this temperature.

   Since the value of Q is much larger than K_{eq}, the reaction must proceed to produce more reactant. The ratio of [products] to [reactants] is much too high. The [products] must decrease and the [reactants] must increase in order to achieve equilibrium. The reaction will proceed in the reverse direction.

Learning Activity 4.3: Interpreting Equilibrium Constants

1. $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2\ HCl_{(g)}$ at 516°C
   A student places 2.00 mol H_2 and 2.00 mol Cl_2 into a 0.500 L container and the reaction is allowed to go to equilibrium at 516°C. If K_{eq} is 76.0, what are the equilibrium concentrations of H_2, Cl_2, and HCl?

2. If $K_{eq} = 1.00 \times 10^{-4}$ for the reaction
   $N_{2(g)} + O_{2(g)} \rightleftharpoons 2\ NO_{(g)}$
   find the number of moles of NO in a 10.0 L container if the initial $[N_2]$ and $[O_2]$ is 1.00 mol/L.

3. If $K_{eq} = 78.0$ for the reaction
   $A_{(s)} + 2\ B_{(g)} \rightleftharpoons 2\ C_{(g)}$
   and initially there are 5.00 moles of A and 4.84 moles of B in a 2.00 L container, how many moles of B are left at equilibrium?

4. For the reaction
   $C_{(s)} + O_{2(g)} \rightleftharpoons CO_{2(g)} \qquad K_{eq} = 25.0$
   find the moles of CO_2 at equilibrium, if initially there are 100.0 moles of C, 50.0 moles of O_2, and 2.0 moles of CO_2 in a 2.00 L container.

5. For the reaction
   $NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)} \qquad K_{eq} = 3.50 \times 10^{-4}$
   find the concentration of NH_3 in a 1.00 L container at equilibrium, if initially there were 0.200 moles of NH_3 added to 0.200 moles of HCl.

6. Equilibrium exists if 5.0 moles of CO_2, 5.0 moles of CO, and 0.20 moles of O_2 are in a 2.0 L container at 562°C.
   $2\ CO_{(g)} + O_{2(g)} \rightleftharpoons 2\ CO_{2(g)}$
   a. Calculate K_{eq} for the reaction at equilibrium.
   b. Would the system be at equilibrium if $[CO_2] = 15.8$ mol/L, $[CO] = 10.0$ mol/L, and $[O_2] = 0.25$ mol/L?

continued

Learning Activity 4.3: Interpreting Equilibrium Constants (continued)

7. For the reaction

   $2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)} \quad K_{eq} = 16.0$

   initially, $[SO_2] = 5.0$ mol/L, $[O_2] = 10.0$ mol/L, and $[SO_3] = 0.$ mol/L. After two hours, $[O_2] = 7.9$ mol/L. Is the system at equilibrium? If not, which substances are increasing and which are decreasing?

8. The reaction

   $4\,NH_{3(g)} + 7\,O_{2(g)} \rightleftharpoons 4\,NO_{2(g)} + 6\,H_2O(g)$

   is at equilibrium when $[H_2O] = 0.100$ mol/L, $[O_2] = 2.00$ mol/L, $[NO_2] = 0.200$ mol/L, and $[NH_3] = 0.500$ mol/L. If 0.75 moles of H_2O, 12.0 moles of NO_2, 30.0 moles of O_2, and 0.30 moles of NH_3 are in a 3.0 L container at the same temperature is equilibrium achieved?

9. Chemists have determined the equilibrium constants for several reactions. In which of these reactions are the products favoured over the reactants?

   a. $K_{eq} = 1.0 \times 10^2$
   b. $K_{eq} = 3.5$
   c. $K_{eq} = 0.003$
   d. $K_{eq} = 6.0 \times 10^{-4}$

Check the answer key.

Lesson Summary

In this lesson, you used the value of the equilibrium constant to explain how far a system at equilibrium has gone toward completion. You also solved different problems involving equilibrium constants. In the next lesson, you will be introduced to Le Châtelier’s principle.

Assignment 4.3: Solving K_{eq} Problems (15 marks)

1. The formation of ammonia from hydrogen and nitrogen occurs as follows:

   $$3\,H_{2(g)} + N_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$$

   Analysis of an equilibrium mixture of nitrogen, hydrogen, and ammonia contained in a 1.0 L flask at 300°C gives the following results: hydrogen, 0.15 moles; nitrogen, 0.25 moles; and ammonia, 0.10 moles. Calculate K_{eq} for the reaction. *(3 marks)*

continued

Assignment 4.3: Solving K_{eq} Problems (continued)

2. If $K_{eq} = 46.0$ for

   $H_{2(g)} + I_{2(g)} \rightleftharpoons 2\ HI_{(g)}$

   what $[I_2]$ would be in equilibrium with 0.50 mol/L HI and 0.10 mol/L H_2?
   (3 marks)

3. Comment on the favourability of product formation in each of the reactions.
   (3 marks)

   a. $H_{2(g)} + F_{2(g)} \rightleftharpoons 2\ HF_{(g)} \quad K_{eq} = 1.0 \times 10^{13}$

   b. $SO_{2(g)} + NO_{2(g)} \rightleftharpoons NO_{(g)} + SO_{3(g)} \quad K_{eq} = 1.0 \times 10^{2}$

   c. $2\ H_2O_{(g)} \rightleftharpoons 2\ H_{2(g)} + O_{2(g)} \quad K_{eq} = 6.0 \times 10^{-28}$

continued

Assignment 4.3: Solving K_{eq} Problems (continued)

4. K_{eq} = 46.0 for the reaction

   $$H_{2(g)} + I_{2(g)} \rightleftharpoons 2\,HI_{(g)}$$

   Initially, there are 6.90 moles of H_2 and 2.40 moles of I_2 in a 1.00 L container. After five hours, there is still 1.00 mole of I_2 left.

   a. Determine the concentration of all species under these conditions. *(3 marks)*

   b. Is the system at equilibrium? If not, which substances are increasing and which are decreasing? *(3 marks)*

Notes

LESSON 4: LE CHÂTELIER'S PRINCIPLE (3 HOURS)

Lesson Focus

SLO C12-4-06: Use Le Châtelier's principle to predict and explain shifts in equilibrium.
Include: temperature changes, pressure/volume changes, changes in reactant/product concentration, the addition of a catalyst, the addition of an inert gas, and the effects of various stresses on the equilibrium constant

Lesson Introduction

In this lesson, you will be introduced to the work of Henri Louis Le Châtelier, a French chemist who investigated how different factors affect the equilibrium of a system. You will then use Le Châtelier's principle to predict and explain some predictable shifts in equilibrium.

Le Châtelier's Principle

Henri Louis Le Châtelier (1850–1936) was a French chemist and a mining engineer. He spent much of his time studying flames in order to prevent mine explosions. He also invented two ways of measuring very high temperatures.

In 1884, Le Châtelier proposed the law of mobile equilibrium, more commonly called **Le Châtelier's principle**. The principle states that if a stress is placed on a **reversible reaction** at **chemical equilibrium**, the equilibrium will shift so as to relieve the stress, thereby restoring equilibrium. Do not forget that this principle only applies to **closed systems**.

Several types of stress can disrupt chemical equilibrium. These stresses include

1. changes in the concentration of reactants or products
2. temperature changes
3. pressure changes
4. volume changes
5. the addition of a catalyst

Generally speaking, chemical systems “like” being at chemical equilibrium; accordingly, if a stress is applied that disrupts this equilibrium, the reaction will shift so as to restore chemical equilibrium. An analogy would be the stress on your body when you are uncomfortably cold. In response, your body starts involuntarily shivering to warm up and you might put on extra clothing. Once you are warm again, your equilibrium (called homeostasis) is returned to its comfortable state. In chemical systems, equilibrium is restored by increasing the rate of either the forward or reverse reactions.

Changes in Concentration

Changing the reactant or product concentrations disturbs equilibrium in the following ways.

- **Adding a reactant** always shifts a reversible reaction in the direction of **products**.
- **Removing a reactant** always shifts a reversible reaction in the direction of **reactants**.
- **Adding a product** causes the reaction to shift in the direction of the formation of the **reactants**.
- **Removing a product** shifts the reaction in the direction of formation of **products**.

Example 1:

Consider the following reversible reaction at equilibrium.

$$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$$

If more reactant (NO_2) is added to the system, the extra reactant must be used up. On the other hand, if more N_2O_4 is added to the system, the extra product must be used up.

According to Le Châtelier’s principle, the position of equilibrium moves to undo the change that you have made. For example, if you decrease the concentration of a product, the position of equilibrium will move to the right to increase the concentration again. **Equilibrium constants aren’t affected by changes in the concentrations of reactants or products at equilibrium.**

If the reaction at equilibrium is a heterogenous system and contains solids, changing the amount of solid will not affect the position of equilibrium.

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

Observe an animation of how the calcium carbonate equilibrium system responds to the addition of a solid at equilibrium at www.dlt.ncssm.edu/tiger/Flash/equilibrium/CaO-CaCO3.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm
2. scrolling down the page until you find the animation titled "**CaO-CaCO3.html**" or "**CaO-CaCO3.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Since a solid or liquid cannot appreciably change in concentration, adding or removing a solid or liquid reactant or product will not affect the position of equilibrium.

Changes in Temperature

Changing the temperature of the system disturbs equilibrium.

Example 2:

Consider the following reaction:

$$2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)} + 20\ kJ$$

Note that this reaction can also be written this way.

$$2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)} \quad \Delta H = -20\ kJ$$

The **enthalpy** change (ΔH) is the amount of heat released or absorbed when a chemical reaction occurs at constant pressure.

This reaction is exothermic, which means that the products release heat (20 kJ) whereas in the reverse reaction, the reactants absorb heat (20 kJ). Alternatively, you could simply remember that a negative enthalpy value denotes an exothermic reaction, and a positive enthalpy value denotes an endothermic reaction. Think of heat as a product; therefore, if the **temperature increases**, the "concentration" of the heat increases. In this example, the equilibrium shifts to the left and the **reactant yield increases**.

On the other hand, if the **temperature decreases**, the equilibrium shifts to the right (the side that releases heat). As a result, the **product yield increases**.

While there are several factors that affect the equilibrium of a reversible reaction, the only factor that changes an equilibrium constant is a change of temperature. Look at the following equilibrium involving hydrogen, iodine, and hydrogen iodide:

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

$$\Delta H = -10.4 \text{ kJ/mol}$$

According to Le Châtelier's principle, the position of equilibrium will move to the left if the temperature increases. Less hydrogen iodide will be formed and the equilibrium mixture will contain more unreacted hydrogen and iodine.

The position of equilibrium changes if you change the temperature. According to Le Châtelier's principle, the position of equilibrium moves in such a way as to undo the change that you have made. If you increase the temperature, the position of equilibrium will move in such a way as to reduce the temperature again. It will do that by favouring the reaction that absorbs heat. You might have already noticed that the ratio of concentrations still stays the same, regardless of the equilibrium shift. This is not the case with reactions involving other reasons for shifting equilibrium (e.g., pressure changes).

Remember that, at equilibrium, where the forward reaction is **exothermic**, increasing the temperature **decreases the value** of the equilibrium constant. If the forward reaction is **endothermic,** increasing the temperature **increases the value** of the equilibrium constant.

Changes in Pressure

A pressure change **only** affects an equilibrium system with an **unequal number of moles** of *gaseous* reactants and products.

- An **increase in pressure** will shift the equilibrium to the side with **fewer moles of gas**.
- A **decrease in pressure** will shift the equilibrium to the side with **more moles of gas**.

Example 3:

Consider the following reaction showing the conversion of dinitrogen monoxide (also called nitrous oxide) to dinitrogen tetroxide. On the reactant side, there are two moles of gas, while on the product side, there is one mole of gas.

$$2\,NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$$

An **increase in pressure** will shift the equilibrium to the side with **fewer moles**, in this case the dinitrogen tetroxide (N_2O_4). An animation to see the shift in equilibrium as the gas is compressed in the vessel available at www.dlt.ncssm.edu/tiger/Flash/equilibrium/2NO2-N204.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm
2. scrolling down the page until you find the animation titled "**2NO2-N204.html**" or "**2NO2-N204.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Before the **increase in pressure**, there are two molecules of N_2O_4 in the vessel. After the pressure increases and the equilibrium shifts to favour the products, there are four molecules of N_2O_4.

Likewise, a **decrease in pressure** will shift the equilibrium to the side with **more moles**; in this case, the reactants. If there are the same number of gaseous moles on each side of the equation, then a change in pressure makes no difference to the position of equilibrium. An animation where you can observe the hydrogen iodide equilibrium is available at www.dlt.ncssm.edu/tiger/Flash/equilibrium/H2I2_equilibrium.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm
2. scrolling down the page until you find the animation titled "**H2I2_equilibrium.html**" or "**H2I2_equilibrium.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

In either scenario, **equilibrium constants aren't affected by changes in the pressure of a system at equilibrium** as long as the reactants and the products are all gases.

Adding a Catalyst

The position of equilibrium *is not* affected by adding a catalyst, since a catalyst speeds up both the forward and reverse reactions by exactly the same amount. For a reaction to proceed at a faster rate without increasing the temperature, increasing the concentration, or increasing the surface area of solid reactants, the activation energy (E_A) must be lowered. Since a catalyst will decrease the E_A in both directions, no shift in the position of equilibrium will result.

Likewise, equilibrium constants are not affected by adding a catalyst since there is no shift in the position of equilibrium.

Learning Activity 4.4: Predicting Shifts in Equilibrium

1. For the reaction

   $PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)} \quad \Delta H = -92.5 \text{ kJ}$

   predict the effect on the position of the equilibrium that results from

   a. increasing the total pressure by decreasing volume
   b. injecting more Cl_2 gas without changing the volume
   c. increasing the temperature
   d. increasing the volume of the container
   e. adding a catalyst

2. For the reaction

   $CH_{4(g)} + H_2O_{(g)} + 49.3 \text{ kJ} \rightleftharpoons CO_{(g)} + 3\, H_{2(g)}$

   predict the effect on the position of the equilibrium that results from

   a. increasing temperature
   b. decreasing temperature
   c. decreasing the pressure
   d. decreasing the volume of the container
   e. adding a solid drying agent such as $CaCl_2$ which reacts with $H_2O_{(g)}$

3. For the reaction

   $9.4 \text{ kJ} + 2\, HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$

   what is the effect on [HI]

   a. if a small amount of H_2 is added?
   b. if the pressure of the system is increased?
   c. if the temperature is increased?
   d. if a catalyst is added?

continued

Learning Activity 4.4: Predicting Shifts in Equilibrium (continued)

4. For the rection

   $CO_{(g)} + 2\,H_{2(g)} \rightleftharpoons CH_3OH_{(g)} + \text{energy}$

   predict the effect of the following changes on the equilibrium concentration of $CH_3OH_{(g)}$.

   a. a decrease in temperature
   b. an increase in pressure
   c. an addition of $H_{2(g)}$
   d. an addition of a catalyst

5. In the equilibrium reaction,

   $2\,NO_{(g)} + O_{2(g)} \rightleftharpoons 2\,NO_{2(g)} + 114.6\ \text{kJ}$

   what will be the change in the equilibrium $[NO_2]$ under each of the following conditions?

   a. O_2 is added
   b. NO is removed
   c. energy is added

6. Match the change to the equilibrium system below with the letter of the appropriate answer. Each letter can be used more than once or not at all.

   $2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)}$

   ___ O_2 is added to the reaction.
   ___ SO_3 is removed from the reaction.
   ___ SO_3 is added from the reaction.
   ___ The pressure of the system is increased.

   a. The equilibrium shifts to the right.
   b. The equilibrium shifts to the left.
   c. There is no change in the equilibrium position.

Check the answer key.

Lesson Summary

In this lesson, you used Le Châtelier’s principle to predict and explain shifts in equilibrium. You also interpreted the effects of various stresses in the equilibrium constant. In the next lesson, you will interpret concentration versus time graphs, including how to identify when equilibrium has been achieved.

Assignment 4.4: Le Châtelier's Principle (14 marks)

For the purpose of this assignment, assume that all reactions occur in a **closed system**.

1. For the following reaction

   $N_2O_{4(g)} + 58.9\text{ kJ} \rightleftharpoons 2\ NO_{2(g)}$

   how will the equilibrium $[NO_2]$ be affected by the following?

   a. an increase in pressure *(1 mark)*

   b. an increase in temperature *(1 mark)*

   c. the addition of a catalyst *(1 mark)*

continued

Assignment 4.4: Le Châtelier's Principle (continued)

2. Suggest four ways to increase the $[SO_3]$ in the following equilibrium reaction. *(4 marks)*

   $2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)} + 192.3\ \text{kJ}$

3. For the reaction

   $4\ HCl_{(g)} + O_{2(g)} \rightleftharpoons 2\ H_2O_{(g)} + 2\ Cl_{2(g)} + 114.4\ \text{kJ}$

   explain how each of the following affects the position of the equilibrium:

   a. the size of the container is increased *(1 mark)*

   b. energy is added *(1 mark)*

   c. $O_{2(g)}$ is added *(1 mark)*

continued

Assignment 4.4: Le Châtelier's Principle (continued)

d. $HCl_{(g)}$ is removed *(1 mark)*

e. a catalyst is added *(1 mark)*

4. Nitric oxide gas (NO) releases 57.3 kJ/mol of energy when it reacts with oxygen gas to yield nitrogen dioxide gas according to the following equation.

   $2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)} + 57.3\text{ kJ}$

   Predict the effect that increasing the temperature will have on

   a. the equilibrium concentrations *(1 mark)*

   b. the speed of formation of NO_2 *(1 mark)*

Notes

LESSON 5: INTERPRETING GRAPHS (2 HOURS)

Lesson Focus

SLO C12-4-08: Interpret concentration-versus-time graphs. Include: temperature changes, concentration changes, and the addition of a catalyst

Lesson Introduction

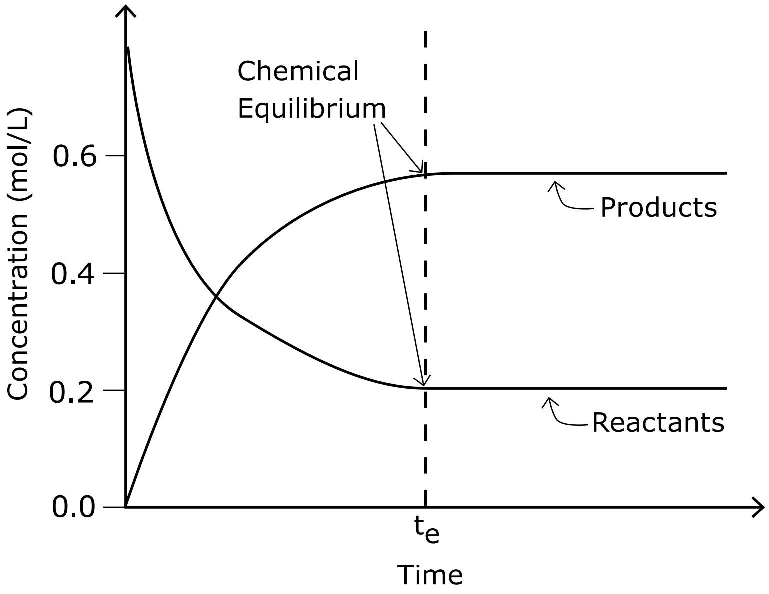
Up to this point, you have learned both theoretical and mathematical aspects of equilibrium. In this lesson, you will interpret equilibrium graphically on a concentration versus time graph.

Interpreting Concentration-versus-Time Graphs

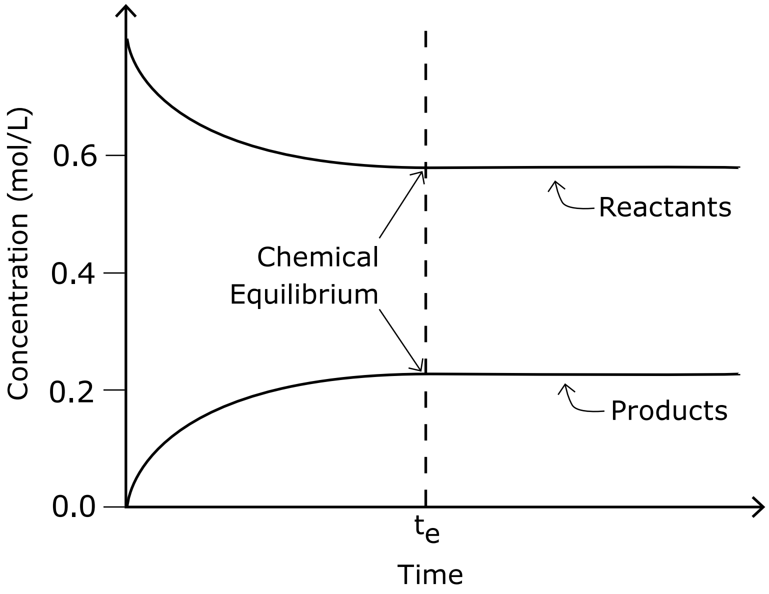
A concentration-versus-time graph has a characteristic pattern as equilibrium is reached. Using the graph provided below as an example, note the following trends.

- The lines representing the concentration of reactants and products will move in the opposite direction. One line will tend upward while the other will tend downward.
- When equilibrium is reached, the lines representing the concentration of reactants and products become flat horizontal lines, also called a plateau. This plateau demonstrates that the concentrations of reactants and products are not changing over time.
- Lines representing concentrations of reactant and products may meet. At this point, the concentration of the reactant and product are the same. This does not mean that the system has reached equilibrium.

In Module 3, we looked at graphs of product and reaction concentrations as they changed over time. Remember that the slope of the tangents to these curves gives us the reaction rate. A plateau, being horizontal, has a slope of 0 indicating that the reaction has stopped. In equilibrium reactions, however, we know that both the forward and reverse reactions are happening all the time. At equilibrium, the rate of the forward reaction equals the rate of the reverse direction, which results in no gain in products or loss of reactants, and a horizontal plateau.



Notice that, at equilibrium, the concentration of products and reactants are **not** necessarily equal, **but** the concentrations of both reactants and products remain constant indefinitely. However, the rate of conversion to products is equal to the rate of conversion to reactants. This is what makes the equilibrium dynamic. On the second graph, you can observe the plateau occurring where equilibrium of the system is achieved.



Concentration-time graphs can be used to obtain data about the reaction at a given time. For example, if you wanted to know the concentration of a product 5 seconds into the reaction, you would do the following:

1. Find 5 s on the *x*-axis.
2. Draw a line directly upward from 5 s to the line corresponding to products.
3. Draw another line connecting this point to the *y*-axis horizontally.
4. Read the concentration value at this point.

Likewise, you could use a concentration-time graph to obtain data about the reaction at a given concentration. If you wanted to know how long the reaction must progress for the concentration of reactants to decrease to 0.06 mol/L, you would do the following:

1. Find 0.06 mol/L on the y-axis.
2. Draw a line horizontally from 0.06 mol/L to the line corresponding to reactants.
3. Draw another line vertically downward connecting this point to the x-axis.
4. Read the time at this point.

By definition, **chemical equilibrium** is the state where the concentrations of reactants and products remain constant over time because the rate of the forward reaction is equal to the rate of the reverse reaction. Remember that this equilibrium can only be established in a closed system.

Learning Activity 4.5: Constructing a Concentration-versus-Time Graph

In this learning activity, you will generate a concentration-versus-time graph on the graph paper provided, based on the information provided in the chart below. The graph will represent the synthesis of $FeSCN^{2+}$. Place time on the x-axis and concentration on the y-axis. Since there are three columns for concentrations, each representing a different set of data, you will have three lines on your graph. Note that a stress is placed on the system at 14 s.

Time	$[SCN^-]$	$[Fe^{3+}]$	$[FeSCN^{2+}]$	$[Fe^{3+}]$ After Stress
0.0	2.5	4.0	0.0	
5.0	1.75	3.25	1.0	
10	1.5	3.0	1.0	
11	1.5	3.0	1.0	
12	1.5	3.0	1.0	
13	1.5	3.0	1.0	
14	1.5	3.0	1.0	7.0
15	1.5		1.0	6.0
20	1.25		1.5	5.0
25	1.25		1.5	5.0
30	1.25		1.5	5.0

continued

Learning Activity 4.5: Constructing a Concentration-versus-Time Graph (continued)

Based on your graph, answer the following questions.

1. Write a balanced equation to represent the reaction.
2. How much time was required for the system to reach equilibrium?
3. Use the equilibrium law to calculate the approximate value of the equilibrium constant from the concentrations at 10 seconds and at 20 seconds.
4. How do the two values from 3 and 4 compare? Why?
5. What was the stress that occurred at 14 seconds?
6. How would the addition of a positive catalyst change the shape of this graph?

Check the answer key.

Lesson Summary

In this lesson, you interpreted concentration-versus-time graphs. In the next lesson, you will conclude your study of Le Châtelier's principle by studying some practical applications of this principle.

Notes

Assignment 4.5: Interpreting a Concentration-versus-Time Graph (10 marks)

In this assignment, you will interpret the concentration-versus-time graph you constructed in the learning activity. Use your constructed graph to answer the following questions.

1. What happens to the concentrations of the reactants (Fe^{3+} and SCN^-) as the reaction proceeds? *(1 mark)*

2. What happens to the concentration of the product ($FeSCN^{2+}$) as the reaction proceeds? *(1 mark)*

3. What indicates that a stress was added to the equilibrium system? *(1 mark)*

continued

Assignment 4.5: Interpreting a Concentration-versus-Time Graph (continued)

4. What is the result of there being more ions of Fe^{3+} in the system? *(2 marks)*

5. What do you think is happening around 20 seconds? Why? *(2 marks)*

6. Use the equilibrium law to calculate the approximate value of the equilibrium constant from the concentrations at 25 seconds. *(2 marks)*

7. When the new equilibrium has been established, what is present in the highest concentration? *(1 mark)*

LESSON 6: APPLICATIONS OF LE CHÂTELIER'S PRINCIPLE (2 HOURS)

Lesson Focus

SLO C12-4-09: Describe practical applications of Le Châtelier's principle.
Examples: Haber process, hemoglobin production at high altitude, carbonated beverages, eyes adjusting to light, blood pH, recharging of batteries, turbocharged/supercharged engines, ester synthesis, weather indicators, arrangement of produce, carbonated beverages in a hen's diet...

Lesson Introduction

You already have a general understanding of Le Châtelier's principle. In this lesson, you will investigate some practical applications, including the Haber process, hemoglobin production at high altitude, carbonated beverages, eyes adjusting to light, blood pH, and the recharging of batteries.

Practical Applications of Le Châtelier's Principle

The following examples serve as an indicator of the importance of Le Châtelier's principle in our lives.

Blood pH

Blood contains dissolved carbonic acid in equilibrium with carbon dioxide and water according to the following reaction.

$$H_2CO_{3(aq)} \rightleftharpoons CO_{2(aq)} + H_2O_{(l)}$$

To keep carbonic acid at safe concentrations in the blood, the CO_2 product is exhaled. The removal of a product causes the forward reaction to be favoured, reducing the amount of carbonic acid to keep blood pH within a safe range.

Chicken Eggs and Soda Pop

Eggshells are made of calcium carbonate, $CaCO_{3(s)}$, which is made from carbon dioxide, a product of cellular respiration.

$$3\ CO_{2(g)} \rightleftharpoons 3\ CO_{2(aq)}$$

(chicken breath)

$$3\ H_2O_{(l)} + 3\ CO_{2(aq)} \rightleftharpoons 3\ H_2CO_{3(aq)}$$

(in the blood)

$$3\ H_2CO_{3(aq)} \rightleftharpoons 3\ H^+_{(aq)} + 3\ HCO_3{}^-{}_{(aq)}$$

(in the blood)

$$3\ HCO_3{}^-{}_{(aq)} \rightleftharpoons 3\ CO_3{}^{2-}{}_{(aq)} + 3\ H^+_{(aq)}$$

$$3\ CO_3{}^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} \rightleftharpoons 3\ CaCO_{3(s)}$$

(in the blood) (eggshell)

The net equation, therefore, would be

$$3\ H_2O_{(l)} + 3\ CO_{2(g)} + Ca^{2+}_{(aq)} \rightleftharpoons 6\ H^+_{(aq)} + 3\ CaCO_{3(s)}$$

When chickens get hot, they pant, and decrease the concentration of carbon dioxide in the blood. To offset the stress, the equilibrium will shift in the reverse direction and decrease the amount of calcium carbonate available to make eggshells. This yields eggs with thin shells that break easily.
It is believed that giving chickens carbonated water to drink will shift equilibrium in the forward direction and minimize the effects of panting on warm days. This could allow farmers to minimize the effects without having to install expensive air conditioning in chicken coops.

Hemoglobin Production and Altitude

In the body, hemoglobin is a complex molecule used to transport oxygen to tissues. (We will use the letters Hb to represent this molecule.)

$$Hb_{(aq)} + O_{2(g)} \rightleftharpoons HbO_{2(aq)}$$

In a place such as Mexico City, where the elevation is 2.3 km, atmospheric pressure and oxygen concentration are low. For a person who is not acclimatized to this, it represents a drop in oxygen concentration. To offset the stress, equilibrium favours the reverse direction. This results in **hypoxia**, a lack of oxygen. Hypoxia can cause headaches, nausea, and extreme fatigue. In serious cases, if treatment is not sought quickly, the victim may slip into a coma and die.

Individuals living at high altitudes for extended periods of time adapt to reduced oxygen concentrations by producing more hemoglobin. This shifts equilibrium to the right once more so that the symptoms of hypoxia disappear. Studies have shown that the Sherpas, long-time residents of the Himalayan mountains, have adapted to high altitude conditions by maintaining high levels of hemoglobin in their blood, sometimes as much as 50% more than individuals living at sea level.

Carbonated Beverages

Soft drinks are carbonated under high pressure to create the following equilibrium system:

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)} + \text{heat}$$

When a bottle of soda is opened, the pressure above the carbon dioxide gas is reduced. The system shifts to the left, the solubility of the carbon dioxide drops, and the carbon dioxide bubbles out of solution. Since the reaction is reversed, heat is taken in, making the bottle cold. If the bottle is left open for long periods of time, the pop will go “flat” due to the reduced pressure.

Increasing the temperature of a pop bottle (i.e., leaving it in a warm car on a summer day) will cause equilibrium to shift in the reverse direction, creating more carbon dioxide gas. This generates a pressure that could potentially cause the pop bottle to burst.

Eyes Adjusting to Light

Photoreceptors, cells that contain the visual pigment rhodopsin, line the inner surface of the eyeball. The rhodopsin is made up of opsin (a protein) and retinene (a pigment). When light strikes a photoreceptor, the energy absorbed changes the shape of the retinene portion of the molecule. This forward reaction takes place very quickly. The shape change signals the optic nerve, which carries information to the brain where it is translated into a visual image.

In the absence of light, the retinene is separated from the opsin. In a dark room, the photoreceptors in your eyes take a few minutes to re-equilibrate to a lower light intensity, as the reverse reaction is slower. Moving into a brightly lit room, the photoreceptors in your eyes again take a few minutes to adjust to their new equilibrium due to the slower reverse reaction.

Supercharged and Turbocharged Engines

In a turbocharged engine, air is compressed and heated. This means that there is a higher concentration (50% more) of warmer oxygen reacting with the gasoline. This favours the production of products, which generates more energy for the car. Note that turbochargers in normal engines work best at higher altitudes, where the air is less dense. The steam created by the reaction of the gasoline and oxygen is used to turn a turbine that runs the air compressor. In a supercharger, a belt runs the compressor.

$$\text{Gasoline} + O_2 \rightleftharpoons CO_2 + H_2O + \text{energy}$$

Rechargeable Batteries

You probably use one or more types of rechargeable batteries in your everyday life. Following are three examples of such batteries, along with the balanced chemical equation for each:

- lead-acid batteries, like the ones used in a car

  $PbO_{2(s)} + Pb_{(s)} + 4\,H^+_{(aq)} + 2\,SO_4{}^{2-}{}_{(aq)} \rightleftharpoons 2\,PbSO_{4(s)} + 2\,H_2O_{(l)} + \text{energy}$
- nickel-cadmium batteries, such as those used in calculators

  $Cd_{(s)} + 2\,NiO(OH)_{(s)} + 2\,H_2O_{(l)} \rightleftharpoons 2\,Ni(OH)_{2(s)} + Cd(OH)_{2(s)} + \text{energy}$
- fuel cells used on spacecrafts and remote weather stations

  $2\,H_{2(g)} + O_{2(g)} \rightleftharpoons 2\,H_2O_{(l)} + \text{energy}$

These types of batteries are recharged through the addition of electrical energy. When energy is added to the system, the reverse reaction is favoured, which produces more reactants.

Learning Activity 4.6: Le Châtelier’s Principle at Work

1. When someone takes your photograph, you may see a “ghost” image of the flash for several minutes after the photo is taken. Explain this phenomenon in terms of the rates of the forward and reverse rhodopsin reactions in the eye.
2. When isopentyl alcohol and acetic acid react, they form the pleasant-smelling compound isopentyl acetate (the essence of banana oil).

   $$C_5H_{11}OH_{(aq)} + CH_3COOH_{(aq)} \rightleftharpoons CH_3COOC_5H_{11(aq)} + H_2O_{(l)}$$

   A student adds a drying agent to remove water in an attempt to increase the yield of banana oil. Is this approach reasonable? Explain.
3. How might Le Châtelier’s principle be useful in the chemical industry? For example, how could you ensure a high yield in the production of fertilizer?

Check the answer key.

Lesson Summary

In this lesson, you learned about several real-life applications of Le Châtelier’s principle. In the next lesson, you will learn about the solubility product expressions and solve some problems relating to salts with low solubilities.

Notes

Assignment 4.6: The Haber Process (16 marks)

For this assignment, you will need to research the Haber process, another example of Le Châtelier's principle at work. You can find the information you need in most chemistry textbooks or on the Internet. The following are three sites that you may find helpful:

- www.ausetute.com.au/haberpro.html
- www.fact-index.com/f/fr/fritz_haber.html
- www.chemguide.co.uk/physical/equilibria/haber.html

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

If you do not have access to the Internet, you will have to find other ways of conducting research. One of the best ways is to go to a school library or public library and ask a librarian for help. Bring your assignment to the librarian so that he or she knows exactly what you are looking for. If you do not have access to a library, look for the information in an encyclopedia or reference book. If you have trouble finding the information, contact your learning partner or tutor/marker.

1. Write the balanced chemical equation, including states, for the Haber process. *(2 marks)*

2. Give two examples of the **uses** of the Haber process. *(2 marks)*

3. What conditions are optimal for the Haber process to proceed? *(2 marks)*

continued

Assignment 4.6: The Haber Process (continued)

4. Describe the history of the Haber process. *(2 marks)*

5. Explain how Karl Bosch used this same process. *(2 marks)*

6. Describe how increasing the pressure of the system will shift the equilibrium and why it will affect the yield of ammonia. *(3 marks)*

7. Describe how decreasing the temperature of the system will shift the equilibrium and why it will affect the yield of ammonia. *(3 marks)*

LESSON 7: SOLUBILITY PRODUCT EXPRESSIONS (3 HOURS)

Lesson Focus

SLO C12-4-10: Write solubility product (K_{sp}) expressions from balanced chemical equations for salts with low solubility.

SLO C12-4-11: Solve problems involving K_{sp}.
Include: common ion problems

Lesson Introduction

Equilibrium also relates to the dissolving of certain salts. At equilibrium, the rate of dissolving is equal to the rate of recrystallization, or the amount of solid is in equilibrium with its ions in solution. In this lesson, you will write solubility product (K_{sp}) expressions from balanced chemical equations for salts with low solubility. You will then use your problem-solving skills to solve problems involving K_{sp}.

Solubility Product Constant (K_{sp})

Ionic compounds, such as salts, differ in their solubilities in water. While a solubility table would indicate that AgCl is insoluble, it does undergo both dissociation and precipitation at equilibrium. Most insoluble salts, such as AgCl, will dissolve to some extent in water and are thus called **slightly soluble**.

At equilibrium, the rate of dissolving is equal to the rate of recrystallization, or the amount of solid is in equilibrium with its ions in solution. Such a solution is said to be **saturated**. Chemical equations for saturated solutions are written as follows:

$$AgCl_{(s)} \rightleftharpoons Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$$

We can refer to this type of equation as the **dissociation equation**.

A good animation of this reaction at the particulate level is available at www.dlt.ncssm.edu/core/Chapter14-Gas_Phase-Solubility-Complex_Ion_Equilibria/Chapter14-Animations/Solubility_of_AgCl.html.

You can also access this animation by

1. going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem5.htm
2. scrolling down the page until you find the animation titled "**Solubility_of_AgCl.html**" or "**Solubility_of_AgCl.exe**"

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

In Lesson 2, you calculated equilibrium constants using the ratio of product concentrations (raised to the value of their coefficients from the balanced equation) to reactant concentrations (raised to the value of their coefficients from the balanced equation) at equilibrium. Based on what you've learned, the equilibrium constant for this equation is as follows:

$$K_{eq} = \frac{\left[Ag^{+}_{(aq)}\right]\left[Cl^{-}_{(aq)}\right]}{\left[AgCl_{(s)}\right]}$$

Since solids are not included in equilibrium expressions because their concentrations are constant, **solubility product constants** are calculated using only the concentrations of products at equilibrium and, because these equilibrium expressions involve the dissolving of a solid, we change K_{eq} to K_{sp}.

$$K_{sp} = [Ag^{+}][Cl^{-}]$$

Note that you do not need states in these expressions. Like equilibrium constants, solubility product constants are specific for only one reaction at a particular temperature. The higher the K_{sp} value, the higher the solubility of the salt.

Examples:

The values given are for 25°C.

calcium phosphate	$K_{sp} = 1.2 \times 10^{-26}$
silver bromide	$K_{sp} = 7.7 \times 10^{-13}$
barium fluoride	$K_{sp} = 1.7 \times 10^{-6}$

Units for K_{sp} are not used because they would vary depending on the powers to which the concentrations are raised, such as mol/L to $(mol/L)^2$ to $(mol/L)^3$.

Interpreting K_{sp} Values

K_{sp} values represent the solubility of a salt in water. The magnitude of the K_{sp} value is an indicator of the solubility of the salt in water.

- The **smaller** the K_{sp} value, the **lower the solubility** of a salt.
- The **larger** the K_{sp} value, the **higher the solubility** of a salt.

Note that K_{sp} values are dependent on temperatures because the solubility of a salt varies with temperature. Some of this might seem familiar from your work with solubility curves in Grade 11 Chemistry.

- If the solubility of a salt **increases** when the **temperature increases**, the magnitude of the $\boldsymbol{K_{sp}}$ will also **increase**.
- If the solubility of a salt **decreases** when the **temperature decreases**, the magnitude of the K_{sp} will also decrease.

Based on the examples above, barium fluoride has a higher solubility than the other two salts, $Ca_3(PO_4)_2$ and AgBr, because BaF_2 has the larger K_{sp} value. Calcium phosphate, $Ca_3(PO_4)_2$, will dissolve very slightly in water due to its very low K_{sp} value.

Solving K_{sp} Problems

One of the more basic K_{sp} problems involves calculating the K_{sp} given the molar solubility of a compound.

Example 1:

The solubility of calcium sulfate, $CaSO_4$, is 4.9×10^{-3} mol/L. Calculate the K_{sp} for $CaSO_4$.

Step 1: Write the dissociation equation for $CaSO_4$. Remember that coefficients can represent many things, including concentrations. In addition, you are not required to write a coefficient that is 1. This equation, therefore, tells us that the concentrations of both Ca^{2+} and SO_4^{2-} ions are equal.

$$CaSO_{4(s)} \rightarrow Ca^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Step 2: Write the solubility product or K_{sp} expression.

$$K_{sp} = [Ca^{2+}][SO_4^{2-}]$$

Step 3: Substitute the molar concentrations of the ions, $Ca^{2+}$and SO_4^{2-}, into the K_{sp} expression and solve.

$$K_{sp} = [4.9 \times 10^{-3} \text{ mol/L}][4.9 \times 10^{-3} \text{ mol/L}]$$

$$K_{sp} = 2.4 \times 10^{-5}$$

You can also use an ICE table to solve for the molar solubility of a slightly soluble salt. Return to Lesson 3 if you need a review of the ICE table method.

Example 2:

Calculate the molar solubility of AgCl in pure water at 25°C. K_{sp} for AgCl is 1.8×10^{-10}.

Step 1: Write the dissociation equation for AgCl.

$$AgCl_{(s)} \rightarrow Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Step 2: Set up an ICE table and fill in the values for the unknown ions.

	$AgCl_{(s)}$	$\rightleftharpoons$	$Ag^{+}_{(aq)}$	$+$	$Cl^{-}_{(aq)}$
I	---		0		0
C	---		$+x$		$+x$
E	---		x		x

Step 3: Write the K_{sp} expression and substitute the known values into the expression.

$$K_{sp} = [Ag^{+}][Cl^{-}]$$

$$1.8 \times 10^{-10} = (x)(x)$$

Step 4: Solve for x.

$$1.8 \times 10^{-10} = x^2$$

$$x = 1.34 \times 10^{-5} \text{ mol/L}$$

The molar solubility of AgCl in pure water at 25°C is 1.34×10^{-5} mol/L.

A third problem type allows you to identify the concentration of ions present at equilibrium. In this case, the K_{sp} value of the slightly soluble salt must be provided.

Example 3:

What is the concentration of lead and chloride ions in a saturated lead (II) chloride solution at 25°C? $K_{sp} = 2.0 \times 10^{-5}$

Step 1: Write the dissociation equation for $PbCl_2$.

$$PbCl_{2(s)} \rightarrow Pb^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

Step 2: Set up an ICE table and fill in the table for the unknown values of the ions, x.

	$PbCl_2$	$\rightarrow Pb^{2+}{}_{(aq)}$ +	$2Cl^{-}{}_{(aq)}$
I	---	0	0
C	---	$+x$	$+2x$
E	---	x	$2x$

Step 3: Write the K_{sp} expression and substitute the known values into the expression.

$$K_{sp} = [Pb^{2+}][Cl^{-}]^2$$

$$2.0 \times 10^{-5} = (x)(2x)^2$$

Step 4: Solve for x.

$$2.0 \times 10^{-5} = (x)(4x^2)$$

$$4x^3 = 2.0 \times 10^{-5}$$

$$x^3 = 5.0 \times 10^{-6}$$

$$x = 1.71 \times 10^{-2}$$

The molar solubilities of the ions at equilibrium are as follows:

$$[Pb^{2+}] = x = 1.71 \times 10^{-2} \text{ mol/L}$$

$$[Cl^{-}] = 2x = 3.42 \times 10^{-2} \text{ mol/L}$$

Note

Do not confuse **solubility** (the number of moles of solute that will dissolve in 1 L of solution [concentration]) with **solubility product** (the product of the concentrations of ions in solution, raised to the powers of their coefficients in the balanced equation).

The Common Ion Effect

In a mixture of salts in solution, a **common ion** is an ion that is common to both salts. As a result of adding a common ion, the ionic compounds become less soluble (for example, when equilibrium is established in a saturated solution of lead (II) chromate and its ions in solution).

rate of dissolving = rate of recrystallization

For the following reaction:

$$PbCrO_{4(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + CrO_4{}^{2-}{}_{(aq)}$$

$$K_{sp} = [Pb^{2+}][CrO_4{}^{2-}] = 1.8 \times 10^{-14}$$

What would happen if $Pb(NO_3)_2$ was added to the solution? Lead nitrate would completely ionize in solution according to the following reaction.

$$Pb(NO_3)_{2(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2\,NO_3{}^{-}{}_{(aq)}$$

You can see that both salts contain the lead ion and, as such, it is the common ion. This would cause the overall $[Pb^{2+}]$ to increase, which in turn would cause the product of $[Pb^{2+}]$ and $[CrO_4{}^{2-}]$ to increase.

The addition of a common ion causes a stress on the system. According to Le Châtelier's principle, the stress on the system will be relieved by shifting the equilibrium position to the left, to use up the additional lead ions.

The lead ions combine with chromate ions to form the original solid lead chromate. $PbCrO_{4(s)}$ will continue to precipitate until the product of the $[Pb^{2+}]$ and $[CrO_4{}^{2-}]$ ions once again equals 1.8×10^{-14}. As a result, the solubility of lead chromate is decreased.

- The solubility of $PbCrO_4$ is less in the presence of $Pb(NO_3)_2$ than it is in pure water.
- Adding $Pb(NO_3)_2$ to $PbCrO_4$ produces the common ion effect.

Finally, you can determine the molar solubility of a slightly soluble salt in a solution containing a known concentration of a common ion.

Example 4:

Calculate the molar solubility of silver chloride in a 1.5×10^{-3} mol/L silver nitrate solution. K_{sp} for AgCl = 1.6×10^{-10}

Solution:

This is a **common ion** problem in which we have two salts in a solution. The common ion is Ag^+, which is present in AgCl and $AgNO_3$. Note that the presence of the common ion affects the solubility of AgCl (in mol/L) but not the K_{sp} value, because it is an equilibrium constant.

Step 1: Write the dissociation equations.

$AgNO_3$ dissociates completely as given by the equation,

$$AgNO_{3(s)} \rightarrow Ag^+_{(aq)} + NO_3{}^-_{(aq)}$$

Since the concentration of $AgNO_3$ is given as 1.5×10^{-3} mol/L, then $[Ag^+] = 1.5 \times 10^{-3}$ mol/L.

The dissociation equation for AgCl is

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

Step 2: Set up an ICE table and fill in the values for the unknown ions. Remember that there are two sources of Ag^+ ions (1.5×10^{-3} mol/L from $AgNO_3$ and the unknown amount, x, from AgCl).

	$AgCl_{(s)}$ →	$Ag^+_{(aq)}$ +	$Cl^-_{(aq)}$
I	---	1.5×10^{-3}	0
C	---	$+ x$	$+x$
E	---	$1.5 \times 10^{-3} + x$	x

Step 3: Write the K_{sp} expression and substitute the known values into the expression.

$$K_{sp} = [Ag^+][Cl^-]$$

$$1.6 \times 10^{-10} = (1.5 \times 10^{-3} + \underset{\uparrow}{x})(x)$$

This x can be ignored because the amount of Ag^+ ion that can dissolve from AgCl is very small compared to the amount of Ag^+ generated from $AgNO_3$.

Step 4: Solve for x.

$$1.6 \times 10^{-10} = (1.5 \times 10^{-3})(x)$$

$$x = 1.1 \times 10^{-7}$$

$$[AgCl] = 1.1 \times 10^{-7} \text{ mol/L}$$

The molar solubility of AgCl in a 1.5×10^{-3} mol/L solution $AgNO_{3(aq)}$ is 1.1×10^{-7} mol/L.

Note

Another commonly used unit is g/L (grams of solute per litre of solution). Solubility units always express the maximum amount of solute that will dissolve in either a given amount of solvent or a given amount of solution, at a specific temperature.

Use the following to convert between units of mol/L and g/L:

$$\begin{tikzcd} \text{mol/L} \arrow[r, bend left, "\text{multiply by molar mass}"] & \text{g/L} \arrow[l, bend left, "\text{divide by molar mass}"] \end{tikzcd}$$

Example 5:

The solubility of copper (II) iodate, $Cu(IO_3)_2$, is 3.3×10^{-3} mol/L. What is the solubility in g/L?

Step 1: Determine the molar mass of copper (II) iodate.

Cu = 63.55 x 1 = 63.55 g/mol

I = 126.9 x 2 = 253.8 g/mol

O = 16.00 x 6 = 96.00 g/mol

Molar mass = 413.35 g/mol

Step 2: Multiply the concentration in mol/L by the molar mass in g/mol.

3.30×10^{-3} mol/L x 413.35 g/mol = 1.36 g/L

Learning Activity 4.7: Solving K_{sp} Problems

1. Write the dissociation equation and the solubility product expression for each of the following (assume that all dissolved solid exists as ions).
   a. $PbSO_4$
   b. $Al_2(SO_4)_3$
   c. $Ba(OH)_2$
   d. $CuCl$
   e. Ag_2CO_3
   f. $Fe_2(SO_4)_3$
2. Given the following compounds' K_{sp}, calculate their solubilities in mol/L and g/L.
   a. CuS $K_{sp} = 6.31 \times 10^{-36}$
   b. PbI_2 $K_{sp} = 1.39 \times 10^{-8}$
   c. SrC_2O_4 $K_{sp} = 1.58 \times 10^{-7}$
3. From the following solubilities, calculate the K_{sp}.
   a. $Pb(OH)_2$ 4.20×10^{-6} mol/L
   b. AgI 2.88×10^{-6} g/L
   c. $Ca_3(PO_4)_2$ 7.15×10^{-7} mol/L
   d. CaF_2 1.70×10^{-2} g/L
4. If 6.7×10^{-5} g of AgBr is all that can be dissolved at 25°C in 500.0 mL, calculate the solubility product of AgBr.
5. A saturated solution of calcium hydroxide has a hydroxide ion concentration of 3.0×10^{-3} mol/L. Calculate the K_{sp} of calcium hydroxide.
6. What are the equilibrium concentrations of all the ions in a saturated solution of AgCN at 25°C, if the K_{sp} is 1.6×10^{-14}?
7. At 25°C, a saturated solution of iron (III) hydroxide has an iron (III) ion concentration of 1.3×10^{-13} mol/L. Calculate the K_{sp} of iron (III) hydroxide.
8. What are the equilibrium concentrations of all the ions in a saturated solution of $Cu(OH)_2$ at 25°C, if the K_{sp} is 1.6×10^{-19}?

Check the answer key.

Lesson Summary

In this lesson, you learned how to write (K_{sp}) expressions from balanced chemical equations for salts with low solubility. You then solved several types of problems involving K_{sp}, including common ion problems. In the next lesson, you will investigate examples of salts with low solubilities.

Assignment 4.7: Working with Solubility Product Constant (20 marks)

1. Write the dissociation equation and the expression for the solubility product constant for $SrSO_4$. *(2 marks)*

2. Write the dissociation equation and the expression for the solubility product constant for $Al_2(SO_4)_3$. *(2 marks)*

For questions 3 to 6, show all the steps of your work. Include correct units and significant digits where they are required.

3. A sample of $Ba(OH)_{2(s)}$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Ba^{2+} is found to be 0.108 mol/L, and that of OH^- is 0.216 mol/L. What is the value of K_{sp} for $Ba(OH)_{2(s)}$? *(4 marks)*

continued

Assignment 4.7: Working with Solubility Product Constant (continued)

4. What is the molar solubility of a saturated solution of AgCl?
$K_{sp} = 1.6 \times 10^{-10}$ *(4 marks)*

5. What will be the equilibrium concentrations of Ca^{2+} and OH^{-} in a saturated solution of $Ca(OH)_2$, if its K_{sp} value is 1.3×10^{-6}? *(4 marks)*

6. Calculate the molar solubility of $Ca(IO_3)_2$ in 0.060 mol/L $NaIO_3$. The K_{sp} of $Ca(IO_3)_2$ is 7.1×10^{-7}. *(4 marks)*

LESSON 8: SALTS WITH LOW SOLUBILITIES (1 HOUR)

Lesson Focus

SLO C12-4-12: Describe examples of salts with low solubility. *Examples: kidney stones, limestone caverns, osteoporosis, tooth decay...*

Lesson Introduction

In the previous lesson, you learned about equilibrium as it pertains to salts with low solubilities. You saw that the addition of a common ion further reduces the solubility of ionic salts. In this lesson, you will read about examples of salts with low solubilities with which you may be familiar, such as limestone caverns and osteoporosis.

Salts with Low Solubilities

The following examples serve as an indicator of the importance of slightly soluble salts in our lives.

Limestone Caverns

Limestone ($CaCO_3$) is formed through the decay of marine organisms like snails, clams, corals, and algae. In water, the slightly soluble salt will set up the following equilibrium:

$$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$$

The chemical erosion of limestone occurs when it is in contact with acidic water.

$$H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow HCO_3^-{}_{(aq)}$$

If the limestone deposit is deep enough underground, the dissolution of the limestone produces a cave.

Osteoporosis

Approximately 99 per cent of the body's calcium is stored in the bones of the skeletal system, where it forms the following equilibrium system:

$$Ca_3(PO_4)_{2(s)} \rightleftharpoons 3\ Ca^{2+}{}_{(aq)} + 2\ PO_4{}^{3-}{}_{(aq)}$$

If the concentration of calcium in the blood decreases, balance can be restored if the dissolving of the calcium phosphate (bone) increases. This leads to porous, brittle bones. How can this be prevented? By making sure to obtain the minimum daily requirement of calcium (especially between the ages of 10 and 20, when bone growth is most rapid), and through regular weight-bearing exercise.

Tooth Decay

The major constituent of tooth enamel is hydroxyapatite ($Ca_5(PO_4)_3OH$, $K_{sp} = 6.8 \times 10^{-37}$). In the mouth, the following equilibrium is established:

$$Ca_5(PO_4)_3OH_{(s)} \rightleftharpoons Ca_5(PO_4)_3{}^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

When sugar ferments on the teeth, the hydronium ion is produced. It reacts with the hydroxide ion from the previous reaction, causing the forward reaction to be favoured. An increase in the solubility of the hydroxyapatite leads to the dissolving of tooth enamel. In recent years, fluoride has been added to water and toothpaste. The fluoride ion replaces the hydroxide ion in hydroxyapatite to create fluorapatite ($Ca_5(PO_4)_3F$, $K_{sp} = 1.0 \times 10^{-60}$). Because the fluorapatite is less soluble in water, teeth become more resistant to cavities.

Note

As helpful as the addition of fluoride to toothpaste has been, it is interesting that fluoride is not added to children's toothpastes. This is because an excess of fluoride in the body from swallowing large amounts of paste can lead to fluorosis, which can cause pigmentary changes to the teeth.

Lesson Summary

In this lesson, you learned about some everyday examples of salts with low solubilities. This is the last lesson in Module 4!

Assignment 4.8: Kidney Stones (5 marks)

For the last assignment of this module, you will conduct research on kidney stones to help you answer the following questions. You may find it helpful to use a search engine such as www.google.ca/ to research kidney stones.

If you do not have access to the Internet, you will have to find other ways of conducting research. One of the best ways is to go to a school library or public library and ask a librarian for help. Bring your assignment to the librarian so that he or she knows exactly what you are looking for. If you do not have access to a library, look for the information in an encyclopedia or reference book. If you have trouble finding the information, contact your learning partner or tutor/marker.

1. Name three examples of salts with low solubilities that may cause kidney stones. *(3 marks)*

2. Name two factors that can lead to the formation of kidney stones. *(2 marks)*

Notes

MODULE 4 SUMMARY

In this module, you examined why and when equilibrium exists, as well as the concentrations of products and reactants at equilibrium. In addition, you determined whether or not a system is at equilibrium and, if not, predicted what the system will do to attain equilibrium.

You have now completed Module 4. You are more than halfway through the course. It is now time to submit your assignments. Read the following directions and double-check that you have submitted all the required work.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 4 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:

☐ Cover Sheet for Module 4 (found at the end of the course Introduction)
☐ Assignment 4.1: Dynamic Equilibrium
☐ Assignment 4.2: Writing Equilibrium Law
☐ Assignment 4.3: Solving K_{eq} Problems
☐ Assignment 4.4: Le Châtelier's Principle
☐ Assignment 4.5: Interpreting a Concentration versus Time Graph
☐ Assignment 4.6: The Haber Process
☐ Assignment 4.7: Working with Solubility Product Constant
☐ Assignment 4.8: Kidney Stones

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

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Notes

GRADE 12 CHEMISTRY (40S)

Module 4
Chemical Equilibrium

Learning Activity Answer Keys

MODULE 4:
CHEMICAL EQUILIBRIUM

Learning Activity 4.1: Equilibrium Terms

1. Match the terms with the correct description.

e	It is the formation of products from reactants.	a. reversible reaction
a	It is a chemical reaction in which the products can return to the original reactants.	b. chemical equilibrium
d	It is the speed at which a reaction occurs.	c. reverse reaction
c	It is the formation of reactants from products.	d. reaction rate
b	It occurs when the rate of formation of products is equal to the rate of formation of reactants over time.	e. forward reaction

2. Does a reaction stop once it has reached equilibrium? Explain.

   No, equilibrium implies that the forward and reverse reactions occur at the same rate.

3. Why is chemical equilibrium referred to as dynamic equilibrium?

   The forward and reverse reactions are both occurring, but with no net change in concentration.

Learning Activity 4.2: The Law of Chemical Equilibrium

1. Match the terms with the correct description below. In some cases, there is more than one correct answer.

f	It is the equilibrium condition for a chemical reaction involving substances in more than one state.	a. equilibrium position
b, c ,g or e	It is used to determine if a reaction has reached equilibrium.	b. law of chemical equilibrium
a	It depends on the initial concentrations of the substances in a reaction.	c. reaction quotient
e or b	It states that every reaction proceeds to an equilibrium state with a specific K_{eq}.	d. homogeneous equilibrium
e	It expresses the relative concentration of reactants and products at equilibrium in terms of an equilibrium constant.	e. law of mass action
d	It is the equilibrium condition for reactions in which products and reactants are in the same state.	f. heterogeneous equilibrium
g	It is the ratio of product concentration to reactant concentration at equilibrium.	g. equilibrium constant

2. Write the equilibrium law (mass action expression) for each of the following reactions.

   a. $SO_{2(g)} + NO_{2(g)} \rightleftharpoons SO_{3(g)} + NO_{(g)}$

   $$K_{eq} = \frac{[\mathbf{SO_3}][\mathbf{NO}]}{[\mathbf{SO_2}][\mathbf{NO_2}]}$$

   b. $2\,C_{(s)} + 3\,H_{2(g)} \rightleftharpoons C_2H_{6(g)}$

   $$K_{eq} = \frac{[\mathbf{C_2H_6}]}{[\mathbf{H_2}]^3}$$

c. $3\ O_{2(g)} \rightleftharpoons 2\ O_{3(g)}$

$$K_{eq} = \frac{[O_3]^2}{[O_2]^3}$$

d. $MgCO_{3(s)} \rightleftharpoons CO_{2(g)} + 2\ MgO_{(s)}$

$$K_{eq} = \frac{[CO_2]}{1}$$

e. $NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$

$$K_{eq} = \frac{[NH_3][HCl]}{1}$$

Learning Activity 4.3: Interpreting Equilibrium Constants

1. $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2\ HCl_{(g)}$ at 516°C
   A student places 2.00 mol H_2 and 2.00 mol Cl_2 into a 0.500 L container and the reaction is allowed to go to equilibrium at 516°C. If K_{eq} is 76.0, what are the equilibrium concentrations of H_2, Cl_2, and HCl?

$$[H_2] = [Cl_2] = \frac{mol}{V} = \frac{2.00\ mol}{0.500\ L} = 4.00\ mol/L$$

	$H_{2(g)}$	$+\ Cl_{2(g)}$	$2\ HCl_{(g)}$
I	**4.00**	**4.00**	**0**
C	**–x**	**–x**	**+ 2x**
E	**4.00 – x**	**4.00 – x**	**2x**

Write the equilibrium law and substitute the equilibrium values.

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$76.0 = \frac{(2x)^2}{(4.00 - x)^2}$$

$$\sqrt{76.0} = \sqrt{\frac{(2x)^2}{(4.00 - x)^2}}$$

$$8.718 = \frac{2x}{4.00 - x}$$

$$(4.00 - x)8.718 = \left(\frac{2x}{4.00 - x}\right)(4.00 - x)$$

$$34.871 - 8.718x = 2x$$

$$34.871 = 2x + 8.718x = 10.718x$$

$$\frac{34.871}{10.718} = x$$

$$3.25 = x$$

Determine the equilibrium concentrations using the value of *x*.

$$[H_2] = [Cl_2] = 4.00 - x = 4.00 - 3.25 = 0.75 \text{ mol/L}$$

$$[HCl] = 2x = 2(3.25) = 6.50 \text{ mol/L}$$

2. If $K_{eq} = 1.00 \times 10^{-4}$ for the reaction
   $N_{2(g)} + O_{2(g)} \rightleftharpoons 2\ NO_{(g)}$
   find the number of moles of NO in a 10.0 L container if the initial $[N_2]$ and $[O_2]$ is 1.00 mol/L.

	$N_{2(g)}$	$+ O_{2(g)}$	$2\ NO_{(g)}$
I	**1.00**	**1.00**	**0**
C	**−x**	**−x**	**+ 2x**
E	**1.00 − x**	**1.00 − x**	**2x**

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$

Write the equilibrium law and substitute the equilibrium values.

$$1.00\times10^{-4} = \frac{(2x)^2}{(1.00-x)^2}$$

$$\sqrt{1.00\times10^{-4}} = \sqrt{\frac{(2x)^2}{(1.00-x)^2}}$$

$$0.0100 = \frac{2x}{1.00-x}$$

$$(1.00-x)0.0100 = \left(\frac{2x}{1.00-x}\right)(1.00-x)$$

$$0.0100 - 0.0100x = 2x$$

$$0.0100 = 2x + 0.0100x = 2.01x$$

$$\frac{0.0100}{2.01} = x$$

$$0.00498 = x$$

$$[NO] = 2x = 2(0.00498) = 0.00996 \text{ mol/L}$$

Determine the equilibrium concentration of NO using the value of *x*.

moles NO = C • V = (0.00996 mol/L)(10.0L) = 0.0996 moles NO

Find the number of moles of NO by multiplying the concentration times the volume.

3. If K = 78.0 for the reaction

$A_{(s)} + 2\,B_{(g)} \rightleftharpoons 2\;C_{(g)}$

and initially there are 5.00 moles of A and 4.84 moles of B in a 2.00 L container, how many moles of B are left at equilibrium?

$$\mathbf{C = \frac{mol}{V} = \frac{4.84\ moles}{2.00\ L} = 2.42\ mol/L\ B}$$

	$A_{(s)}$	$+\ 2\ B_{(g)}$	$2\ C_{(g)}$
I	--	**2.42**	**0**
C	--	**−x**	**+ x**
E	--	**2.42 − x**	**x**

$$K_{eq} = \frac{[\mathrm{C}]^2}{[\mathrm{B}]^2}$$

Write the equilibrium law and substitute the equilibrium values.

$$78.0 = \frac{x^2}{(2.42 - x)^2}$$

$$\sqrt{78.0} = \sqrt{\frac{(2x)^2}{(2.42 - x)^2}}$$

$$8.832 = \frac{x}{2.42 - x}$$

$$(2.42 - x)8.832 = \left(\frac{x}{2.42 - x}\right)(2.42 - x)$$

$$21.37 - 8.832x = x$$

$$21.37 = 9.832x$$

$$\frac{21.37}{9.832} = x$$

$$2.17 = x$$

Determine the equilibrium concentration of B using the value of *x*.

$$[\mathrm{B}] = 2.42 - x = 2.42 - 2.17 = 0.25\ \mathrm{mol/L}$$

Find the number of moles of B by multiplying the concentration times the volume.

moles B = CV = (0.25 mol/L) (2.00 L) = 0.50 moles B

4. For the reaction

$C_{(s)} + O_{2(g)} \rightleftharpoons CO_{2(g)} \qquad K_{eq} = 25.0$

Find the moles of CO_2 at equilibrium, if initially there are 100.0 moles of C, 50.0 moles of O_2, and 2.0 moles of CO_2 in a 2.00 L container.

$$\mathbf{C_{O_2} = \frac{mol}{v} = \frac{50.0\ moles}{2.00\ L} = 25.0\ mol/L}$$

$$\mathbf{C_{CO_2} = \frac{mol}{v} = \frac{2.0\ moles}{2.00\ L} = 1.00\ mol/L}$$

	$C_{(s)}$	$+\ O_{2(g)}$	$CO_{2(g)}$
I	**--**	**25.0**	**1.00**
C	**--**	**$-x$**	**$+ x$**
E	**--**	**$25.0 - x$**	**$1.00 + x$**

$$K_{eq} = \frac{[CO_2]}{[O_2]}$$

Write the equilibrium law and substitute the equilibrium values.

$$25.0 = \frac{1.00 + x}{25.0 - x}$$

Eliminate the denominator by multiplying both sides by the denominator.

$$25.0(25.0 - x) = \left(\frac{1.00 + x}{25.0 - x}\right)(25.0 - x)$$

$$625 - 25.0x = 1.00 + x$$

$$624 = 26x$$

Solve for x, then for the equilibrium concentration of CO_2.

$$\frac{624}{26} = \frac{26x}{26}$$

$$24.0 = x$$

$$[CO_2] = 1.00 + x = 25.0\ \text{mol/L}$$

Find the moles of CO_2.

$$\text{moles} = CV = (25.0\ \text{mol/L})(2.00\ \text{L}) = 50.0\ \text{moles}\ CO_2$$

5. For the reaction

   $NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)} \qquad K_{eq} = 3.50 \text{ x } 10^{-4}$

   Find the concentration of NH_3 in a 1.00 L container at equilibrium if initially there were 0.200 moles of NH_3 added to 0.200 moles of HCl.

	$NH_4Cl_{(s)}$	$NH_{3(g)}$	$+ HCl_{(g)}$
I	--	**0.200**	**0.200**
C	--	**−x**	**− x**
E	--	**0.200 − x**	**0.200 − x**

$$K_{eq} = [NH_3][HCl]$$

$$3.50 \text{ x } 10^{-4} = \left(0.200 - x^2\right)$$

$$\sqrt{3.50 \text{ x } 10^{-4}} = \sqrt{\left(0.200 - x^2\right)}$$

$$0.01871 = 0.200 - x$$

$$0.181 = x$$

$$[NH_3] = 0.200 - x = 0.200 - 0.181$$

$$[NH_3] = 0.019 \text{ mol/L } [NH_3]$$

6. Equilibrium exists if 5.0 moles of CO_2, 5.0 moles of CO, and 0.20 moles of O_2 are in a 2.0 L container at 562°C.

   $2\,CO_{(g)} + O_{2(g)} \rightleftharpoons 2\,CO_{2(g)}$

   a. Calculate K_{eq} for the reaction at equilibrium.

$$[CO_2] = [CO] = \frac{5.0 \text{ mol}}{2.0 \text{ L}} = 2.5 \text{ mol/L}$$

$$[O_2] = \frac{0.20 \text{ mol}}{2.0 \text{ L}} = 0.10 \text{ mol/L}$$

$$K_{eq} = \frac{[CO_2]^2}{[CO]^2[O_2]}$$

$$K_{eq} = \frac{(2.5)^2}{(2.5)^2(0.10)}$$

$$K_{eq} = 10$$

$$Q = \frac{[CO_2]^2}{[CO]^2[O_2]} = \frac{(15.8)^2}{(10.0)^2(0.25)} \approx 10$$

b. Would the system be at equilibrium if $[CO_2] = 15.8$ mol/L, $[CO] = 10.0$ mol/L, and $[O_2] = 0.25$ mol/L?

Since $Q = K_{eq}$, the system is at equilibrium.

7. For the reaction

$$2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)} \qquad K_{eq} = 16.0$$

Initially, $[SO_2] = 5.0$ mol/L, $[O_2] = 10.0$ mol/L, and $[SO_3] = 0$ mol/L. After two hours, $[O_2] = 7.9$ mol/L. Is the system at equilibrium? If not, which substances are increasing and which are decreasing?

	$2\,SO_{2(g)}$	$+ O_{2(g)}$	$\rightleftharpoons 2\,SO_{3(g)}$
Initial	**5.0**	**10.0**	**0.0**
Change	**−4.2**	**−2.1**	**+4.2**
Final	**0.8**	**7.9**	**4.2**

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(4.2)^2}{(0.8)^2(7.9)} = 3.5$$

Q is less than K_{eq}; therefore, the reaction must proceed to produce more product. This means that the $[SO_2]$ and $[O_2]$ are decreasing and $[SO_3]$ is increasing.

8. The reaction

$$4\,NH_{3(g)} + 7\,O_{2(g)} \rightleftharpoons 4\,NO_{2(g)} + 6\,H_2O(g)$$

is at equilibrium when $[H_2O] = 0.100$ mol/L, $[O_2] = 2.00$ mol/L, $[NO_2] = 0.200$ mol/L, and $[NH_3] = 0.500$ mol/L. If 0.75 moles of H_2O, 12.0 moles of NO_2, 30.0 moles of O_2, and 0.30 moles of NH_3 are in a 3.0 L container at the same temperature is equilibrium achieved?

First, find K_{eq},

$$K_{eq} = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7} = \frac{(0.200)^4(0.100)^6}{(0.500)^4(2.00)^7} = 8.00 \times 10^{-10}$$

Then, find Q,

$$Q = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7} = \frac{(4.0)^4(0.25)^6}{(0.10)^4(10.0)^7} = 6.25 \times 10^{-3}$$

9. Chemists have determined the equilibrium constants for several reactions. In which of these reactions are the products favoured over the reactants?

   a. $K_{eq} = 1.0 \times 10^2$
   b. $K_{eq} = 3.5$
   c. $K_{eq} = 0.003$
   d. $K_{eq} = 6.0 \times 10^{-4}$

   A and B will favour product formation because the K_{eq} values are larger than one.

Learning Activity 4.4: Predicting Shifts in Equilibrium

1. For the reaction

   $PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)} \quad \Delta H = -92.5\ kJ$

   predict the effect on the position of the equilibrium that results from

   a. increasing the total pressure by decreasing volume

      Decreasing the volume of the container increases the total pressure on the system. According to Le Châtelier's principle, the system will lower the pressure by reducing the number of particles. This causes the equilibrium to shift right (toward products) because the right side of the equation has fewer particles.

   b. injecting more Cl_2 gas without changing the volume

      Adding chlorine gas to the system increases the concentration of a reactant. According to Le Châtelier's principle, the system will use up the added reactant by shifting toward the products (right).

   c. increasing the temperature

      According to Le Châtelier's principle, increasing the temperature tends to favour the endothermic reaction. This system will respond by shifting the equilibrium to the left (reactants) in order to use the added heat.

   d. increasing the volume of the container

      Increasing the volume of the container reduces the pressure of the system. According to Le Châtelier's principle, the system will respond by increasing the pressure. The system will shift to the left (reactants), since the reactants have more particles.

   e. adding a catalyst

      A catalyst lowers the activation energy of both forward and reverse reactions equally. Adding a catalyst does not affect the equilibrium position, just the speed with which equilibrium is reached.

Note: Question 1 describes how each answer should be written as a full answer. The following questions will only provide the answers, not the explanations.

2. For the reaction

   $CH_{4(g)} + H_2O_{(g)} + 49.3\ kJ \rightleftharpoons CO_{(g)} + 3\ H_{2(g)}$

   predict the effect on the position of the equilibrium that results from

   a. increasing temperature

      Increasing the temperature favours the endothermic reaction; therefore, the equilibrium shifts right (toward products).

   b. decreasing temperature

      Decreasing the temperature causes a shift to the left (reactants), in order to replace the lost heat.

   c. decreasing the pressure

      Decreasing the pressure shifts the equilibrium to the right (products).

   d. decreasing the volume of the container

      Decreasing the volume of the container increases the pressure, shifting the equilibrium to the left. This increases the pressure by increasing particles.

   e. adding a solid drying agent such as $CaCl_2$ which reacts with $H_2O_{(g)}$

      Adding a solid drying agent such as $CaCl_2$, which reacts with $H_2O_{(g)}$, reduces water concentration. Since water is a reactant, the system will shift toward the left (reactants) to replace the lost reactant.

3. For the reaction

   $9.4\ kJ + 2\ HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$

   what is the effect on [HI]

   a. if a small amount of H_2 is added?

      increases [HI]

   b. if the pressure of the system is increased?

      no change in [HI] since both sides have two particles each

   c. if the temperature is increased?

      decreases [HI]

   d. if a catalyst is added?

      no change in [HI] if a catalyst is added

4. For the reaction

   $CO_{(g)} + 2\,H_{2(g)} \leftrightarrow CH_3OH_{(g)} + \text{energy}$

   predict the effect of the following changes on the equilibrium concentration of $CH_3OH_{(g)}$.

   a. a decrease in temperature

      increases [CH_3OH]

   b. an increase in pressure

      increases [CH_3OH]

   c. an addition of $H_{2(g)}$

      increases [CH_3OH]

   d. an addition of a catalyst

      does not affect the equilibrium of [CH_3OH]

5. In the equilibrium reaction,

   $2\,NO_{(g)} + O_{2(g)} \rightleftharpoons 2\,NO_{2(g)} + 114.6\text{ kJ}$

   what will be the change in the equilibrium [NO_2] under each of the following conditions?

   a. O_2 is added

      [NO_2] increases to use up added oxygen

   b. NO is removed

      [NO_2] decreases to replace lost NO

   c. energy is added

      [NO_2] is decreased to use up added heat

6. Match the change to the equilibrium system below with the letter of the appropriate answer. Each letter can be used more than once or not at all.

   $2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)}$

   a O_2 is added to the reaction.

   a SO_3 is removed from the reaction.

   b SO_3 is added from the reaction.

   a The pressure of the system is increased.

   a. The equilibrium shifts to the right.

   b. The equilibrium shifts to the left.

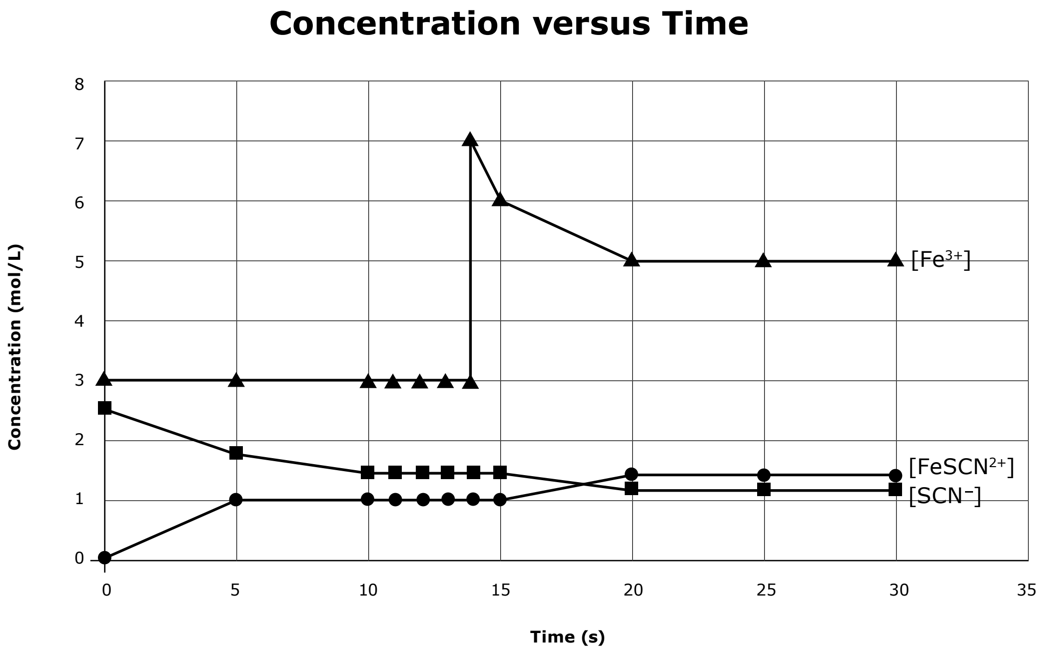
   c. There is no change in the equilibrium position.

Learning Activity 4.5: Constructing a Concentration-versus-Time Graph

In this learning activity, you will generate a concentration-versus-time graph on the graph paper provided, based on the information provided in the chart below. The graph will represent the synthesis of $FeSCN^{2+}$. Place time on the x-axis and concentration on the y-axis. Since there are three columns for concentrations, each representing a different set of data, you will have three lines on your graph. Note that a stress is placed on the system at 14 s.

Time	$[SCN^-]$	$[Fe^{3+}]$	$[FeSCN^{2+}]$	$[Fe^{3+}]$ After Stress
0.0	2.5	4.0	0.0	
5.0	1.75	3.25	1.0	
10	1.5	3.0	1.0	
11	1.5	3.0	1.0	
12	1.5	3.0	1.0	
13	1.5	3.0	1.0	
14	1.5	3.0	1.0	7.0
15	1.5		1.0	6.0
20	1.25		1.5	5.0
25	1.25		1.5	5.0
30	1.25		1.5	5.0

Your graph should look like this:



Based on your graph, answer the following questions.

1. Write a balanced equation to represent the reaction.

   $\mathbf{Fe^{3+} + SCN^{-} \rightleftharpoons FeSCN^{2+}}$

2. How much time was required for the system to reach equilibrium?

   10 seconds

3. Use the equilibrium law to calculate the approximate value of the equilibrium constant from the concentrations at 10 seconds and at 20 seconds.

   $$\textbf{At 10s, } K_{\mathbf{eq}} = \frac{[\mathrm{FeSCN}^{2+}]}{[\mathrm{Fe}^{3+}][\mathrm{SCN}^{-}]} = \frac{1.0}{(1.5)(3.0)} = 0.22$$

   $$\textbf{At 20s, } K_{\mathbf{eq}} = \frac{[\mathrm{FeSCN}^{2+}]}{[\mathrm{Fe}^{3+}][\mathrm{SCN}^{-}]} = \frac{1.5}{(1.25)(5.0)} = 0.24$$

4. How do the two values from 3 and 4 compare? Why?

   They are approximately the same because the stress imposed on the system was not a change in temperature.

5. What was the stress that occurred at 14 seconds?

   It was the addition of $\mathbf{Fe^{3+}}$.

6. How would the addition of a positive catalyst change the shape of this graph?

   A catalyst would decrease the time required to reach equilibrium. This would condense (squish) the graph along the x-axis. The concentrations do not change.

Learning Activity 4.6: Le Châtelier's Principle at Work

1. When someone takes your photograph, you may see a "ghost" image of the flash for several minutes after the photo is taken. Explain this phenomenon in terms of the rates of the forward and reverse rhodopsin reactions in the eye.

   When the flash occurred, the photoreceptors in the eye responded quickly to the bright burst; however, since the reverse reaction is much slower and the intensity of the flash was so great, a ghost image can be seen for several minutes while the reactions in the photoreceptors take time to reverse themselves.

2. When isopentyl alcohol and acetic acid react, they form the pleasant-smelling compound isopentyl acetate (the essence of banana oil).

   $C_5H_{11}OH_{(aq)} + CH_3COOH_{(aq)} \rightleftharpoons CH_3COOC_5H_{11(aq)} + H_2O_{(l)}$

   A student adds a drying agent to remove water in an attempt to increase the yield of banana oil. Is this approach reasonable? Explain.

   Adding a drying agent will decrease the amount of water present in the system. To minimize the stress and re-establish equilibrium, the system will favour the production of more products. Thus, adding a drying agent is a reasonable course of action to increase the yield of banana oil.

3. How might Le Châtelier's principle be useful in the chemical industry? For example, how could you ensure a high yield in the production of fertilizer?

   Le Châtelier's principle could be used to maximize the amount of product formed in chemical reactions. By altering temperature and pressure, constantly adding reactants, and immediately removing products, the products continue to be favoured.

Learning Activity 4.7: Solving K_{sp} Problems

1. Write the dissociation equation and the solubility product expression for each of the following (assume that all dissolved solid exists as ions).

   a. $PbSO_4$

   $\mathbf{PbSO_{4(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \qquad K_{sp} = [Pb^{2+}][SO_4{}^{2-}]}$

   b. $Al_2(SO_4)_3$

   $\mathbf{Al_2(SO_4)^3{}_{(s)} \rightleftharpoons 2\ Al^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)} \qquad K_{sp} = [Al^{3+}]^2[SO_4{}^{2-}]^3}$

   c. $Ba(OH)_2$

   $\mathbf{Ba(OH)_{2(s)} \rightleftharpoons Ba^{2+}{}_{(aq)} + 2\ OH^-{}_{(aq)} \qquad K_{sp} = [Ba^{2+}][OH^-]^2}$

   d. $CuCl$

   $\mathbf{CuCl_{(s)} \rightleftharpoons Cu^+{}_{(aq)} + Cl^-{}_{(aq)} \qquad K_{sp} = [Cu^+][Cl^-]}$

   e. Ag_2CO_3

   $\mathbf{Ag_2CO_{3(s)} \rightleftharpoons 2\ Ag^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \qquad K_{sp} = [Ag^+]^2[CO_3{}^{2-}]}$

   f. $Fe_2(SO_4)_3$

   $\mathbf{Fe_2(SO_4)_{3(s)} \rightleftharpoons 2\ Fe^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)} \qquad K_{sp} = [Fe^{3+}]^2[SO_4{}^{2-}]^3}$

2. Given the following compounds' K_{sp}, calculate their solubilities in mol/L and g/L.

   a. CuS K_{sp} = 6.31 x 10^{-36}

	$CuS_{(s)}$	$\rightleftharpoons Cu^{2+}{}_{(aq)}$	$+ S^{2-}{}_{(aq)}$
I	--	**0**	**0**
C	--	**+ x**	**+ x**
E	--	**x**	**x**

$$\mathbf{K_{sp} = \left[Cu^{+}\right]\left[S^{2-}\right]}$$

$$\mathbf{6.31 \times 10^{-36} = x^2}$$

$$\mathbf{\sqrt{6.31 \times 10^{-36}} = \sqrt{x^2}}$$

$$\mathbf{2.51 \times 10^{-18} = x}$$

The amount of CuS dissolved will be equal to the [Cu^{2+}] or x.

The concentration of the saturated solution is 2.51 x 10^{-18} mol/L.

The mass of CuS in 1 L of a saturated solution is
mass = (mol)(g/mol) = (2.51 x 10^{-18} mol)(95.6 g/mol) = 2.40 x 10^{-16} g

The solubility of CuS is 2.51 x 10^{-18} mol/L or 2.40 x 10^{-16} g/L.

   b. PbI_2 K_{sp} = 1.39 x 10^{-8}

	$PbI_{2(s)}$	$\leftrightarrow Pb^{2+}{}_{(aq)}$	$+ I^{-}{}_{(aq)}$
I	--	**0**	**0**
C	--	**+ x**	**+ 2x**
E	--	**x**	**2x**

$$\mathbf{K_{sp} = \left[Pb^{2+}\right]\left[I^{-}\right]^2}$$

$$\mathbf{1.39 \times 10^{-8} = x(2x)^2}$$

$$\mathbf{\frac{1.39 \times 10^{-8}}{4} = \frac{4x^3}{4}}$$

$$\mathbf{3.475 \times 10^{-9} = x^3}$$

$$\mathbf{\sqrt[3]{3.475 \times 10^{-9}} = \sqrt[3]{x^3}}$$

$$\mathbf{1.51 \times 10^{-3} = x}$$

Dissolved PbI_2 = [Pb^{2+}] = x

The concentration of the saturated solution is 1.5 x 10^{-3} mol/L.

The mass of PbI_2 in 1 L of a saturated solution is
mass = (mol)(g/mol) = (1.51 x 10^{-3} mol)(461.0 g/mol) = 0.696 g.

The solubility of PbI_2 is 1.51 x 10^{-3} mol/L or 0.696 g/L.

c. SrC_2O_4 $K_{sp} = 1.58 \times 10^{-7}$

	$SrC_2O_{4(s)}$	$\leftrightarrow Sr^{2+}{}_{(aq)}$	$+ C_2O_4{}^{2-}{}_{(aq)}$
I	--	**0**	**0**
C	--	**+ x**	**+ x**
E	--	**x**	**x**

$$K_{sp} = \left[Sr^{2+}\right]\left[C_2O_4{}^{2-}\right]$$

$$1.58 \times 10^{-7} = x^2$$

$$\sqrt{1.58 \times 10^{-7}} = \sqrt{x^2}$$

$$3.97 \times 10^{-4} = x$$

Dissolved SrC_2O_4 = $[Sr^{2+}]$ = x

The concentration of the saturated solution is 3.97×10^{-4} mol/L.

The mass of SrC_2O_4 in 1 L of a saturated solution is
mass = (mol)(g/mol) = (3.97×10^{-4} mol)(175.6 g/mol) = 6.97×10^{-2} g.

The solubility of SrC_2O_4 is 3.97×10^{-4} mol/L or 6.97×10^{-2} g/L.

3. From the following solubilities, calculate the K_{sp}.

a. $Pb(OH)_2$ 4.20×10^{-6} mol/L

$$Pb(OH)_{2(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2\,OH^{-}{}_{(aq)}$$

From the stoichiometry,

$$[Pb^{2+}] = [Pb(OH)_2] = 4.20 \times 10^{-6} \text{ mol/L}$$

$$[OH^-] = 2[Pb^{2+}] = 8.40 \times 10^{-6} \text{ mol/L}$$

$$K_{sp} = [Pb^{2+}][OH^-]^2$$

$$= (4.20 \times 10^{-6})(8.40 \times 10^{-6})^2$$

$$K_{sp} = 2.96 \times 10^{-16}$$

b. AgI 2.88 x 10^{-6} g/L

$\mathbf{AgI_{(s)} \rightleftharpoons Ag^{+}{}_{(aq)} + I^{-}{}_{(aq)}}$

Determine the concentration in mol/L:

$$\text{moles AgI} = \frac{\text{mass}}{\text{g/mol}} = \frac{2.88 \text{ x } 10^{-6}\,\text{g}}{234.8 \text{ g/mol}} = 1.226 \times 10^{-8} \text{ moles AgI}$$

$$\left[\mathrm{Ag}^{+}\right] = \left[\mathrm{I}^{-}\right] = \left[\mathrm{AgI}\right] = 1.226 \times 10^{-8} \text{ mol/L}$$

$$K_{\mathrm{sp}} = \left[\mathrm{Ag}^{+}\right]\left[\mathrm{I}^{-}\right]$$

$$K_{\mathrm{sp}} = \left(1.226 \times 10^{-8}\right)\left(1.226 \times 10^{-8}\right)$$

$$K_{\mathrm{sp}} = 1.50 \times 10^{-16}$$

c. $\mathrm{Ca_3(PO_4)_2}$ 7.15 x 10^{-7} mol/L

$\mathbf{Ca_3(PO_4)_{2(s)} \rightleftharpoons 3\,Ca^{2+}{}_{(aq)} + 2\,PO_4{}^{3-}{}_{(aq)}}$

From the stoichiometry,

$[\mathrm{Ca}^{2+}] = 3 \text{ x } [\mathrm{Ca}^{3}(\mathrm{PO}_4)^{2}] = 2.145 \text{ x } 10^{-6}$

$[\mathrm{PO_4}^{3-}] = 2 \text{ x } [\mathrm{Ca}^{3}(\mathrm{PO}_4)^{2}] = 1.43 \text{ x } 10^{-6}$

$$\begin{aligned} K_{\mathrm{sp}} &= [\mathrm{Ca2+}]^{3}[\mathrm{PO_4}^{3-}]^{2} \\ &= (2.145 \text{ x } 10^{-6})^{3}(1.43 \text{ x } 10^{-6})^{2} \\ K_{\mathrm{sp}} &= 2.02 \text{ x } 10^{-29} \end{aligned}$$

d. $\mathrm{CaF_2}$ 1.70 x 10^{-2} g/L

$\mathbf{CaF_{2(s)} \rightleftharpoons Ca^{2+}{}_{(aq)} + 2\,F^{-}{}_{(aq)}}$

Determine the concentration in mol/L:

$$\text{moles CaF}_2 = \frac{\text{mass}}{\text{g/mol}} = \frac{1.70 \times 10^{-2}\text{g}}{78.1 \text{ g/mol}} = 2.177 \times 10^{-4} \text{moles CaF}_2$$

From the stoichiometry,

$[\mathrm{Ca}^{2+}] = [\mathrm{CaF_2}] = 2.177 \text{ x } 10^{-4}$ mol/L

$[\mathrm{F}^{-}] = 2 \text{ x } [\mathrm{Ca}^{2+}] = 4.35 \text{ x } 10^{-4}$

$$\begin{aligned} K_{\mathrm{sp}} &= [\mathrm{Ca}^{2+}][\mathrm{F}^{-}]^{2} \\ &= (2.177 \text{ x } 10^{-4})(4.35 \text{ x } 10^{-4})^{2} \\ K_{\mathrm{sp}} &= 4.12 \text{ x } 10^{-11} \end{aligned}$$

4. If 6.7 x 10^{-5} g of AgBr is all that can be dissolved at 25°C in 500.0 mL, calculate the solubility product of AgBr.

Determine the concentration of the solution:

$$\text{moles AgBr} = \frac{\text{mass}}{\text{g/mol}} = \frac{6.7 \times 10^{-5}\text{g}}{187.8\ \text{g/mol}} = 3.57 \times 10^{-7}\text{moles AgBr}$$

$$\text{g/mol} = \frac{\text{mol}}{\text{v}} = \frac{3.57 \times 10^{-7}\text{moles}}{0.500\ \text{L}} = 7.14 \times 10^{-7}\text{moles AgBr}$$

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

Substitute into solubility product expression:

$K_{sp} = [Ag^+][Br^-]$

$= (7.14 \times 10^{-7})(7.14 \times 10^{-7})$

$K_{sp} = 5.1 \times 10^{-13}$

5. A saturated solution of calcium hydroxide has a hydroxide ion concentration of 3.0 x 10^{-3} mol/L. Calculate the K_{sp} of calcium hydroxide.

$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

$[OH^-] = 3.0 \times 10^{-3}$

$[Ca^{2+}] = \frac{1}{2}[OH^-] = 1.5 \times 10^{-3}$

Substitute into solubility product expression:

$K_{sp} = [Ca^{2+}][OH^-]^2$

$= (1.5 \times 10^{-3})(3.0 \times 10^{-3})^2$

$K_{sp} = 1.35 \times 10^{-8}$

6. What are the equilibrium concentrations of all the ions in a saturated solution of AgCN at 25°C, if the K_{sp} is 1.6 x 10^{-14}?

	$AgCN_{(s)}$	$\rightleftharpoons Ag^+_{(aq)}$	$+ CN^-_{(aq)}$
I	--	**0**	**0**
C	--	**+ x**	**+ x**
E	--	**x**	**x**

$K_{sp} = [Ag^+][CN^-]$

$1.6 \times 10^{-14} = x^2$

$\sqrt{1.6 \times 10^{-14}} = \sqrt{x^2}$

$1.26 \times 10^{-7} = x$

7. At 25°C, a saturated solution of iron (III) hydroxide has an iron (III) ion concentration of 1.3 x 10^{-13} mol/L. Calculate the K_{sp} of iron (III) hydroxide.

   $Fe(OH)_{3(s)} \rightleftharpoons Fe^{3+}{}_{(aq)} + 3\ OH^{-}{}_{(aq)}$

   From stoichiometry,

   $[OH^-] = 3 \times [Fe^{3+}] = 3.9 \times 10^{-13}$ mol/L

   Substitute into solubility product expression:

   $$K_{sp} = [Fe^{3+}][OH^-]^3$$
   $$= (1.3 \times 10^{-13}\ \text{mol/L})(3.9 \times 10^{-13}\ \text{mol/L})^3$$
   $$K_{sp} = 7.7 \times 10^{-51}$$

8. What are the equilibrium concentrations of all the ions in a saturated solution of $Cu(OH)_2$ at 25°C, if the K_{sp} is 1.6 x 10^{-19}?

	$Cu(OH)_{2(s)}$	$\rightleftharpoons Cu^{2+}{}_{(aq)}$	$+ 2\ OH^{-}{}_{(aq)}$
I	--	**0**	**0**
C	--	**+ x**	**+ 2x**
E	--	**x**	**2x**

$$K_{sp} = [Cu^{2+}][OH^-]^2$$
$$1.6 \times 10^{-19} = x(2x)^2$$
$$\frac{1.6 \times 10^{-19}}{4} = \frac{4x^3}{4}$$
$$4.0 \times 10^{-20} = x^3$$
$$\sqrt[3]{4.0 \times 10^{-20}} = \sqrt[3]{x^3}$$
$$3.4 \times 10^{-7} = x$$
$$[Cu^{2+}] = x = 3.4 \times 10^{-7}\ \text{mol/L}$$
$$[OH^-] = 2x = 6.8 \times 10^{-7}\ \text{mol/L}$$

GRADE 12 CHEMISTRY (40S)

Module 5
Acids and Bases

Note to Students

Module 5 contains an image that is best viewed in colour. A colour version in PDF format is available in the learning management system (LMS).

MODULE 5:
ACIDS AND BASES

Introduction to Module 5

Acids and bases are two important classes of chemical compounds. You encounter acids and bases every day in foods and items such as personal care products (for example, soap and toothpaste), medications, and many others. In addition, acids and bases are very important to the manufacturing industry in the production of dyes and textiles, paper, polymers, and fertilizers, just to name a few.

There are acids and bases at work in your body too. For example, body systems are very sensitive to the acidity of your blood. Your stomach secretes hydrochloric acid to aid in digestion. You might occasionally take an antacid to control the amount of acid in your stomach.

Acids and bases play a large part in our everyday lives. In this module, you will define acids and bases and study acid-base reactions.

Note

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.

Assignments in Module 5

Note

Please remember to work on Assignment 1.9: Researching a Redox Application as you work through this course. Please submit it at the end of Module 6.

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	Acid-Base Theories
2	Assignment 5.2	Identifying Conjugate Pairs
3	Assignment 5.3	Solving K_w Problems
4	Assignment 5.4	Solving pH Problems
5	Assignment 5.5	Acid and Base Equilibrium
6	Assignment 5.6	pH and Percent Dissociation
7	Assignment 5.7	Titration Curves
8	Assignment 5.8	Acid-Base Predictions

LESSON 1: ACID-BASE THEORIES (1 hour)

Lesson Focus

SLO C12-5-01: Outline the historical development of acid-base theories.
Include: the Arrhenius, Brønsted-Lowry, Lewis theories

Lesson Introduction

You may remember learning about acids and bases in your Grade 10 Science course. More specifically, you may remember a single definition or theory of acids and bases. In this lesson, we will study the historical development of our knowledge of acids and bases, as well as learn some new definitions of acid-base systems. These theories explain why some substances exhibit acidic or basic properties.

Early Acid-Base Theories

The word **acid** comes from the Latin word *acere*, which means sour. Vinegar and lemons were long known by the ancients to taste sour. The word **alkaline** (another, older word for **base**) was derived from the Arabic word *al-qily*, meaning "roasted in a pan" or "the ashes of plants." Writings as early as the tenth century have shown that mixing the ashes of plants with water and then with slaked lime would produce the alkaline solutions NaOH and KOH.

Acids and bases were initially defined based on observable properties, also called **operational definitions**. In 1661, **Robert Boyle** (1627–1691) described acids as

- sour tasting
- able to change the colour of a vegetable litmus dye red
- corrosive
- becoming less acidic when mixed with alkaline (basic) substances

He described the properties of bases, or alkalines, as

- having a slippery feel
- being able to change litmus blue
- becoming less alkali when mixed with acids

Antoine Lavoisier (1743–1794) created the word oxygen in 1777 after the Greek words *oxys* for sour and *genes* for born. He said that oxygen was an acid-maker. He studied several acids, including nitric and sulphuric acids, and proposed that oxygen was responsible for compounds being acidic. Although he was wrong, he was the first to identify acids chemically instead of by observable properties.

In 1810, **Sir Humphry Davy** began questioning Lavoisier's theory. He noticed that reactions of several acids with metals and non-metals never resulted in the production of oxygen. He proposed that hydrogen gives a compound acidic qualities. Davy's theory was not complete, since many hydrogen-containing compounds, such as ammonia (NH_3), are not acidic.

In 1838, **Justus Liebig** (1803–1873), a German chemist, modified Davy's theory by proposing that acids contain hydrogen, which can be replaced by metals to form a salt. This theory held for about 50 years until Svante Arrhenius' description of acids and bases.

The Arrhenius Theory

In the early 1880s, a Swedish chemist named **Svante Arrhenius** (1859–1927) was studying the conductivity of solutions. At that time, very little was known about the structure of atoms and molecules. He proposed that electrolytes break up into charged particles in water. Scientists of the time did not accept Arrhenius' theory. They believed that atoms and molecules were indestructible and could not break apart.

With the help of some prominent scientists, Arrhenius continued to encourage the scientific community to accept his theory. Eventually, as knowledge of atomic structure increased, Arrhenius' theory was accepted. In 1903, Arrhenius received the Nobel Prize in science for his work.

In 1887, Arrhenius discovered that all acidic and basic solutions he tested were **electrolytes**. He determined that acids and bases must ionize or dissociate in water. This theory is useful when working with compounds whose formulas contain obvious clues about their acidic or basic nature.

Note

According to Arrhenius, an acid is defined as a hydrogen-containing compound that ionizes to yield hydrogen ions (H^+) dissolved in aqueous solution.

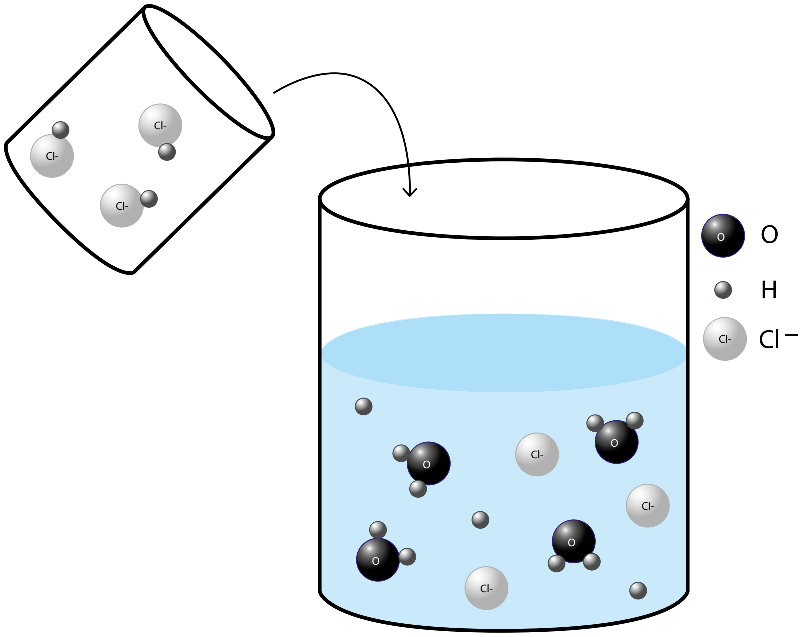
He also stated that a base is defined as a compound that ionizes to yield hydroxide ions (OH^-) in aqueous solution.

Example 1

Hydrochloric acid dissociates in water according to the reaction below.

$$HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

The illustration below shows HCl forming ions in water from a particle view.



Example 2

Acetic acid (vinegar) dissociates in water according to the reaction below.

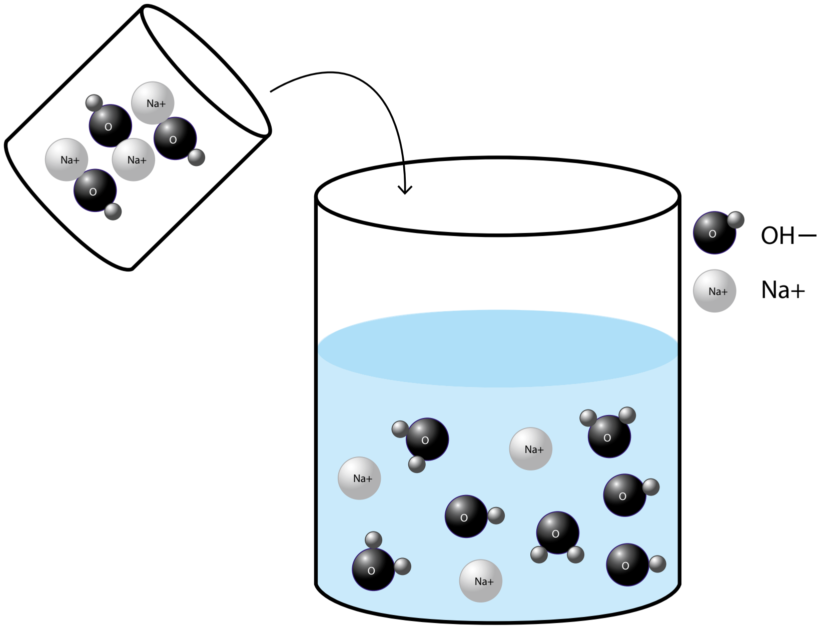
$$HC_2H_3O_{2(aq)} \rightarrow H^{+}_{(aq)} + C_2H_3O_2^{-}{}_{(aq)}$$

Example 3

Sodium hydroxide, a base, dissociates in water according to the reaction below.

$$NaOH_{(s)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

The illustration below shows NaOH forming ions in water from a particulate view.



Until Arrhenius' theory, there was no satisfactory explanation for why acids and bases neutralize each other. He proposed that hydrogen ions produced by acids react with hydroxide ions released by bases to form water.

Limitations of the Arrhenius Theory

According to Arrhenius' theory, acids and bases can only occur in aqueous (water) solutions; however, the theory does not account for reactions between substances that are acidic or basic but do not have a hydrogen ion or a hydroxide ion. Such examples include

- an aqueous solution of baking soda (sodium bicarbonate, $NaHCO_3$), which turns litmus blue in water, but has no apparent hydroxide ion
- metal ions, such as iron (III) and aluminum, which turn litmus red, but have no hydrogen ions

The Arrhenius theory was primarily important in establishing the concept of dissociation and explaining the process of neutralization. It was also an essential starting point for the further development of our understanding of acids and bases.

The Brønsted-Lowry Theory

The limitations of the Arrhenius theory prompted further study by a Danish chemist, **Johannes Brønsted** (1879–1947), and an English chemist, **Thomas Lowry** (1874–1936). These chemists worked independently of one another to develop a more general definition of acids and bases, which they accomplished within months of each other in 1923. Their definition accounted for the acidity of metal ions **and** the basicity (alkalinity) of baking soda. Since they developed essentially the same theory such a short time apart, both their names are applied to this theory of acids and bases.

Note

According to the Brønsted-Lowry definition, an acid is a proton (H^+) donor. A base is defined as a proton (H^+) acceptor.

We call the hydrogen ion (H^+) a proton because, as an ion, it has no electron and the nucleus is a single proton.

Example 4

In the following example, HCl (hydrochloric acid) reacts with water by donating a proton. Water acts as the base, accepting the proton. The result is the H_3O^+ ion called the **hydronium ion**.

$$\overset{\text{H}^+\ \longrightarrow}{\text{HCl}_{(g)} + \text{H}_2\text{O}_{(l)}} \rightarrow \text{Cl}^-_{(aq)} + \text{H}_3\text{O}^+_{(aq)}$$

Acid	**Base**
Proton	Proton
Donor	Acceptor

Example 5

In the next example, ammonia accepts a proton from water, making ammonia a base and water the acid. The result is an ammonium ion (NH_4^+) and a hydroxide ion (OH^-).

$$\overset{\longleftarrow\ \text{H}^+}{\text{NH}_{3(g)} + \text{H}_2\text{O}_{(l)}} \rightarrow \text{OH}^-_{(aq)} + \text{NH}_4^+{}_{(aq)}$$

Base	**Acid**
Proton	Proton
Acceptor	Donor

The following link will take you to an animation showing NH_3 and H_2O combining to form NH_4^+ and OH^-. It also illustrates the Lewis structures for this equilibrium. The reverse reaction is also shown.

www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animationsindex.htm

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

According to the Brønsted-Lowry theory, substances like water can act as both an acid and a base. These types of substances are referred to as **amphoteric** (the root amph is similar to the root of amphibian, which means "having two lives," on land and in water).

Another example of an amphoteric substance is the hydrogen carbonate ion, HCO_3^-, found in baking soda.

The hydrogen carbonate can act as an acid by losing its hydrogen in water

$$HCO_3^- + H_2O \rightarrow CO_3^{2-} + H_3O^+$$

or as a base by stealing a proton from water

$$HCO_3^- + H_2O \rightarrow H_2CO_3 + OH^-$$

Advantages and Problems with the Brønsted-Lowry Theory

The Arrhenius theory of acids and bases was limited to aqueous solutions. The Brønsted-Lowry theory expands the definition of an acid and base to a proton donor or acceptor. However, even though acids and bases can occur without water, there is still a requirement for the presence of a solvent.

The Brønsted-Lowry theory explains why substances such as the hydrogen carbonate ion can act as an acid and a base and how substances without hydroxides, like ammonia, can act as bases. The theory does not explain, however, how substances without protons can act as acids.

Furthermore, Brønsted and Lowry explained acids and bases in terms of proton donors and acceptors. More accurately, bases do not simply accept protons, they take protons away from acids.

The Lewis Theory

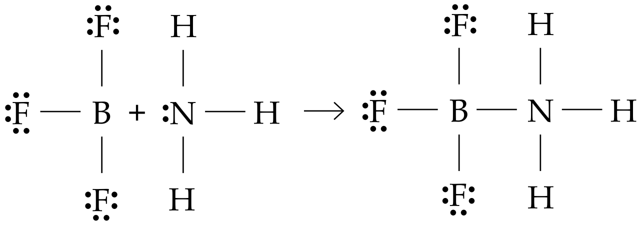
A third, more general, theory of acids and bases was proposed by Gilbert Lewis (1875–1946). Lewis proposed that an acid accepts a pair of electrons during a chemical reaction, while a base donates a pair of electrons.

Note

A **Lewis acid** is a substance that can accept a pair of electrons to form a covalent bond. A **Lewis base** is a substance that donates a pair of electrons to form a covalent bond.

Example 6

According to the Lewis theory, a species that has no hydrogen to donate can still be an acid. As an example, consider the molecule BF_3 (boron trifluoride). Looking at the Lewis structure of BF_3, you can see that boron does not have a complete octet and can, therefore, accept a lone pair of electrons. On doing so, BF_3 acts as a Lewis acid. When reacting with ammonia (NH_3), the reaction would proceed as follows:



Characteristics of the Lewis Theory

- The Lewis definition is so general that any reaction in which a pair of electrons is transferred becomes an acid-base reaction.
- Lewis acid-base reactions include many reactions that would not be included with the Brønsted-Lowry definition.
- The Lewis acid-base theory does not affect the substances we have previously called Brønsted-Lowry bases because any Brønsted-Lowry base must have a pair of non-bonding electrons in order to accept a proton.
- However, this theory vastly expands the number of compounds we have called acids. Any compound that has one or more valence shell orbitals can now act as an acid.

Since many Lewis acid-base reactions can be quite complex, the main focus of this module will be the Brønsted-Lowry acid-base reactions.

Learning Activity 5.1: Acid-Base Definitions

1. Compare the three definitions of acids and bases that were introduced in this lesson.

Theory	Acid	Base
Arrhenius		
Brønsted-Lowry		
Lewis		

2. Match the terms with the correct description below.

_____	1.	It is a substance that can accept a pair of electrons.	a. Amphoteric
_____	2.	It describes a substance that can act as both an acid and a base.	b. Hydroxide ion
_____	3.	OH^-	c. Hydronium ion
_____	4.	H_3O^+	d. Lewis acid
_____	5.	It is the formula of the hydrogen ion.	e. H^+

3. What is the Brønsted-Lowry definition of an acid? Of a base?
4. Identify the Lewis acid and the Lewis base in the following reaction:

   $H^+ + Cl^- \rightarrow HCl$

Check the answer key.

Lesson Summary

In this lesson, you learned about the historical development of acid-base theories. You now have a general knowledge of the Arrhenius, Brønsted-Lowry, and Lewis theories. In the next lesson, you will put your knowledge of chemical equations into practice by writing balanced acid/base chemical equations.

Notes

Assignment 5.1: Acid-Base Theories (5 marks)

1. Provide one explanation as to why NH_3 might **not** be considered a base according to Arrhenius. *(1 mark)*

2. Why is the Brønsted-Lowry definition of acids and bases considered to be more general than the Arrhenius definition? *(2 marks)*

3. Of the three acid-base theories you studied in this lesson, which one applies to the greatest number of reactions? Explain your reasoning. *(2 marks)*

Notes

LESSON 2: ACID-BASE EQUATIONS
(2 HOURS)

Lesson Focus

SLO C12-5-02: Write balanced acid/base chemical equations.
Include: conjugate acid/base pairs, amphoteric behaviour

Lesson Introduction

Now that you know about the three main acid-base theories, you will go one step further and learn to write balanced acid/base chemical equations. In writing these equations, you will also learn to identify conjugate acid and conjugate base pairs.

Dissociation and Ionization

You might recall from your study of solutions that electrolytic solutions are those that conduct electricity because the substance dissolves in water to produce ions. When ionic compounds dissolve to produce ions, the process is typically called **dissociation**.

Dissociation of ionic compounds occurs when water molecules separate the ionic crystal into positive and negative ions. Water molecules then surround the positive cations and negative anions; this is called **hydration**. The following are some examples of dissociation equations:

$$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$Na_2SO_{4(s)} \rightarrow 2\ Na^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$(NH_4)_3PO_{4\ (s)} \rightarrow 3\ NH_4^{+}{}_{(aq)} + PO_4^{3-}{}_{(aq)}$$

$$KOH_{(s)} \rightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$$

$$Mg(OH)_{2\ (s)} \rightarrow Mg^{2+}_{(aq)} + 2\ OH^{-}_{(aq)}$$

When writing dissociation equations, remember that **ion charges** must be included and balanced. An ion, such as the sodium ion Na^{+} is not the same as a sodium atom, Na. Be sure to write the charges for all ions and refer to your **Table of Common Ions** (Appendix H) when needed. You are now familiar with the charges of ions you use most often, including some polyatomic ions.

When molecular compounds dissolve in water to produce ions, the process is typically called **ionization**. Recall that molecular compounds are held together by covalent bonds, and ionic compounds are held together by ionic bonds. While most molecular compounds do not undergo ionization, acids are an exception. Acids produce hydrogen ions in solution. The following are some examples:

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2SO_{4(g)} \rightarrow 2\,H^+_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

$$HC_2H_3O_{2(l)} \rightleftharpoons H^+_{(aq)} + C_2H_3O_2{}^-{}_{(aq)}$$

The reason for using a double arrow for acetic acid (in the last example) will be discussed later in this lesson.

The reactions, as written above, are actually simplified versions of what really occurs. In fact, the hydrogen ion bonds to a water molecule to form the **hydronium ion**, H_3O^+. As a result, we could also write these reactions as follows:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2SO_{4\,(g)} + H_2O_{(l)} \rightarrow 2\,H_3O^+_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

$$HC_2H_3O_{2\,(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O_2{}^-{}_{(aq)}$$

Dissociation Equations

In Grade 11 Chemistry, you learned that when an ionic compound dissolves in water, it no longer exists as a compound, but as freely moving ions that are generally not associated with each other. You will use these concepts again when writing dissociation equations for acids and bases. Here is an overview of the steps of the process, including how to make sure the overall charge is conserved.

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.

Mg^{2+} and Cl^{-}

Step 3: Use the subscripts from the formula to indicate the coefficients to balance the charges.

One magnesium ion will be required to cancel the charge of two chlorine ions. In the chemical equation, you do not need to write the coefficient if it is 1.

$1\ Mg^{2+}$ and $2\ Cl^{-}$

Step 4: Write the equation using the appropriate state for the compound (s) and (aq) for each dissociated ion.

$$MgCl_{2(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2\ Cl^{-}{}_{(aq)}$$

When magnesium chloride is dissolved in water, the compound dissociates into freely moving magnesium ions and chloride ions. Following is another example using an ionic compound.

Example 2

Write the equation for solid aluminum sulphate, $Al_2(SO_4)_3$, dissolved in water.

Step 1: Determine if the compound is ionic or molecular.

Aluminum sulphate is composed of metal (Al) and non-metal (S and O) 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. Remember to separate the metals from the non-metals and look up the corresponding charges.

Al^{3+} and $SO_4{}^{2-}$

Step 3: Use the subscripts from the formula to indicate the coefficients.

Two aluminum ions are required to cancel the charge of three sulfate ions.

$2\ Al^{3+}$ and $3\ SO_4{}^{2-}$

Step 4: Write the equation using the appropriate state for the compound (s) and (aq) for each dissociated ion.

$$Al_2(SO_4)_{3(s)} \rightarrow 2\ Al^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)}$$

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.

Step 2: Write the equation using the appropriate state for the compound (l) and (aq) for the dissolved molecule.

$$CH_3OH_{(l)} \rightarrow CH_3OH_{(aq)}$$

Acids and Bases: A Review

In the previous lesson, you learned that acids are substances that increase the hydronium (H_3O^+) ion concentration. Thus, acids are **proton donors**.

Example 1: hydrochloric acid

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

Example 2: nitric acid

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow NO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

Therefore, when HCl or HNO_3 are added to water, the hydronium ion concentration is increased. Hence they are acids.

You were also introduced to the definition of bases as substances that increase the hydroxide (OH-) ion concentration. Thus, bases are **proton acceptors**.

Example 3: sodium hydroxide

$$NaOH_{(aq)} \rightarrow Na^+{}_{(aq)} + OH^-{}_{(aq)}$$

Example 4: ammonia

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

When NaOH or NH_3 are added to water, the hydroxide ion concentration is increased. Hence they are considered bases.

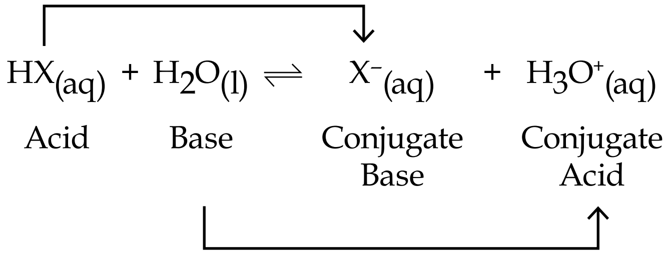
Conjugate Acids and Bases

As mentioned in the previous lesson, the focus will be on the Brønsted-Lowry model of acids and bases. The general form of a Brønsted-Lowry acid-base reaction is

$$\text{Acid} + \text{Base} \rightleftharpoons \text{Conjugate Acid} + \text{Conjugate Base}$$

> **Note**
>
> The **conjugate acid** is what remains after a base has accepted a proton, and the **conjugate base** is what remains after the acid has donated its proton.

In any acid-base reaction, a conjugate acid and base pair are established.



$$HX_{(aq)} + H_2O_{(l)} \rightleftharpoons X^-_{(aq)} + H_3O^+_{(aq)}$$

You can think of it in this way: when an acid donates a proton, the part of the acid left behind can accept a proton in the reverse reaction. Similarly, when a base accepts a proton, the resulting molecule can donate that proton in the reverse reaction.

Example 5:

Let's take another look at the reaction of ammonia with water.

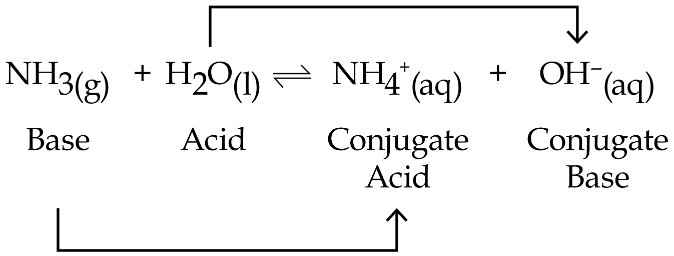
$$\underset{\text{Base}}{NH_{3(g)}} + \underset{\text{Acid}}{H_2O_{(l)}} \rightleftharpoons OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

The reverse reaction would be

$$\underset{\text{Acid}}{NH_4^+{}_{(aq)}} + \underset{\text{Base}}{OH^-_{(aq)}} \rightleftharpoons H_2O_{(l)} + NH_{3(g)}$$

You can see that an acid results from the ammonia accepting a proton from water. The ammonium ion can donate a proton to the hydroxide ion. The hydroxide ion accepts the proton, making it a base.

In the first reaction, ammonia is the base and the ammonium ion is its conjugate acid. Water is the acid in the first reaction and the hydroxide ion is its conjugate base. $NH_{3(g)}$ and $NH_4^+{}_{(aq)}$ are called a **conjugate acid-base pair**, as are $H_2O_{(l)}$ and $OH^-{}_{(aq)}$. This reversible reaction can be written as follows:



$$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Example 6:

Identify the acid, base, conjugate acid, and conjugate base in the following reversible reaction.

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + F^-{}_{(aq)}$$

In this reaction, a proton is transferred from HF to H_2O. When the HF (acid) gives up its proton, what remains is called the conjugate base of that acid. Consequently, F- is a conjugate base because it can now accept a proton.

That makes H_2O a base in this reaction because it accepts a proton. After it has accepted a proton, the water becomes a hydronium ion. It now has a proton (which it can donate); therefore, it is a conjugate acid.

Acids are classified by the number of hydrogen ions available to be donated. **Monoprotic** acids have one hydrogen ion to donate, such as HCl. **Polyprotic** acids have two or more hydrogen ions to donate (di-, tri-, and so on). All polyprotic acids, such as H_2SO_4, can donate one hydrogen ion at a time.

Amphoteric Substances

Most of the reactions you will be studying in this module will be in an aqueous environment. In order to study these reactions further, you need to understand the amphoteric nature of water. This means that water can act as both an acid and a base.

Ionization of Water

Since water can act as an acid or a base, it can accept or donate an H^+ ion. In this first reaction, HA acts as a Brønsted-Lowry acid in water.

$$HA + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+_{(aq)}$$

The HA donates an H^+ ion to water; therefore, it is an acid. The water accepts the H^+ ion, thus acting as a base.

In the next reaction, B acts as a Brønsted-Lowry base in water.

$$B + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}$$

In this case, water acts as the acid by donating an H^+ ion to B. Since B accepts the H^+ ion, it is the Brønsted-Lowry base.

Note

Substances (such as water) that can act as both acids and bases are said to be amphoteric.

Learning Activity 5.2: Writing Balanced Acid-Base Reactions

1. Complete the following acid-base reactions, using the examples provided in the lesson to help you predict the products and assure conservation of charges. Then, identify which of these are Arrhenius acids and bases, and which are Brønsted-Lowry acids and bases.
   a. $HI_{(aq)} + H_2O_{(l)} \rightarrow$

   b. $HF_{(aq)} + H_2O_{(l)} \rightarrow$

   c. $C_2H_3O_2{}^-{}_{(aq)} + H_2O_{(l)} \rightarrow$

   d. $CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow$

   e. $HSO_3{}^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons$

2. Since Brønsted-Lowry acids are proton donors, what do you think the term *triprotic acid* means?
3. What does the Brønsted-Lowry acid-base theory state in regards to water?

Check the answer key.

Lesson Summary

In this lesson, you learned that a conjugate acid is what remains after a base has accepted a proton, and a conjugate base is what remains after the acid has donated its proton. You used this knowledge to help you complete balanced acid-base reactions. In the next lesson, you will describe the relationship between the hydronium and hydroxide ion concentrations in water.

Assignment 5.2: Identifying Conjugate Pairs (14 marks)

1. Identify the **acid, base, conjugate acid**, and **conjugate base** for each of the following. *(2 marks each x 7 = 14 marks)*

   a. $H_2SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HSO_3^-{}_{(aq)} + H_3O^+{}_{(aq)}$

   ____________ ____________ ____________ ____________

   b. $HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$

   ____________ ____________ ____________ ____________

   c. $H_2S_{(g)} + H_2O_{(l)} \rightleftharpoons HS^-{}_{(aq)} + H_3O^+{}_{(aq)}$

   ____________ ____________ ____________ ____________

   d. $HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons SO_3^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$

   ____________ ____________ ____________ ____________

   e. $NH_{3(g)} + H_2O_{(l)} \rightleftharpoons OH^-{}_{(aq)} + NH_4^+{}_{(aq)}$

   ____________ ____________ ____________ ____________

   f. $HF_{(aq)} + HSO_3^-{}_{(aq)} \rightleftharpoons F^-{}_{(aq)} + H_2SO_{3(aq)}$

   ____________ ____________ ____________ ____________

   g. $HNO_{2(aq)} + HS^-{}_{(aq)} \rightleftharpoons NO_2^-{}_{(aq)} + H_2S_{(aq)}$

   ____________ ____________ ____________ ____________

Notes

LESSON 3: THE ION PRODUCT OF WATER (3 HOURS)

Lesson Focus

SLO C12-5-03: Describe the relationship between the hydronium and hydroxide ion concentrations in water.
Include: the ion product of water, K_w

Lesson Introduction

In the previous lesson, you learned about the amphoteric property of water. In this lesson, you will study the relationship between the hydronium and hydroxide ion concentrations in water.

Equilibrium Constant

Remember that many of the solutions you studied in Grade 12 Chemistry involved water as the solvent. In the previous lesson, you revisited the nature of water and learned about its ability to act as an acid or as a base. In pure water, some of the water molecules will dissociate to form ions. This represents a small degree of ionization, as only two molecules out of one billion will ionize.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

This is also called the self-ionization of water.

Since both H_3O^+ and OH^- ions are present, you can see how water could have an effect on both acids and bases. If you take the dissociation equation for water, you can write the following equilibrium expression:

$$K_{eq} = [H_3O^+][OH^-]$$

Note that the concentration of H_2O, $[H_2O]$, is not included because all pure liquids or solids have a constant concentration. Since water is a liquid, the product of K_{eq} and the concentration of water results in the ion product for water, K_w. (This important equilibrium constant is given its own symbol.)

$$K_w = [H_3O^+][OH^-]$$

At 25°C, the concentration of the hydronium and hydroxide ions are equal at 1.0×10^{-7} mol/L. By substituting these values into the expression

$$K_W = [H_3O^+][OH^-]$$

we get

$$K_W = [1 \times 10^{-7}][1 \times 10^{-7}] = 1 \times 10^{-14}$$

Therefore, at 25°C, the value of K_W is 1.0×10^{-14}. Note that this constant, like others you have used, has an infinite number of significant figures. This means you do not consider them when determining the correct number of significant figures for your final answer.

Effect of Adding Acid and Base to Water

Using the ionization of water equation, you can predict the effect of dissolving an acid or base on hydronium and hydroxide ion concentrations by using Le Châtelier's principle. You can see that hydronium (hydrogen) ions **and** hydroxide ions are **both** present in any solution—whether they are acidic or basic.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + OH^-{}_{(aq)}$$

Adding Acid

When an acid is dissolved in water, the acid produces a large amount of H_3O^+ ions. If the H_3O^+ ion concentration increases, the equilibrium will shift to the left to use up some of the added hydronium and maintain K_W at 1.0×10^{-14}. Since equilibrium shifts left, the hydroxide ion concentration is reduced.

Note

Adding an acid to water increases the hydronium ion concentration and reduces the hydroxide ion concentration.

Adding Base

When a base is dissolved in water, the hydroxide ion concentration increases. According to Le Châtelier's principle, the equilibrium shifts left to use up some of the added hydroxide and maintain K_W at 1.0×10^{-14}. Since equilibrium shifts left, the hydronium ion concentration is reduced.

Note

Adding a base to water increases the hydroxide ion concentration and reduces the hydronium ion concentration.

Calculating Ion Concentrations

Using the K_W expression, you can solve for an unknown concentration of hydroxide ions or hydronium ions.

Example 1

If 2.5 moles of hydrochloric acid is dissolved in 5.0 litres of water, what is the concentration of the hydroxide ions? Note that the reaction indicates a 1:1 ratio for all species. Assume the volume remains unchanged.

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$$

Step 1: If 2.5 moles is dissolved in 5.0 litres, the concentration of the HCl would be as follows:

$$\text{Concentration} = \frac{2.5 \text{ moles}}{5.0 \text{ L}} = 0.50 \text{ mol/L HCl}$$

Since HCl is a strong acid and completely dissociates, the $[H_3O^+] = 0.50$ mol/L.

Step 2: Use the equilibrium law for water to find the concentration of the hydroxide ion.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$K_W = [H_3O^+][OH^-]$$

$$\left[OH^-\right] = \frac{K_W}{\left[H_3O^+\right]} = \frac{1.0 \, x \, 10^{-14}}{0.50 \text{ mol/L}} = 2.0 \, x \, 10^{-14} \text{ mol/L } OH^-$$

Example 2

0.40 g of NaOH is dissolved in water to make a solution with a volume of 1.0 L. What is the hydronium ion concentration in this solution?

$$NaOH_{(aq)} + H_2O_{(l)} \rightarrow NaO^-_{(aq)} + H_3O^+_{(aq)}$$

Step 1: Determine the concentration of the hydroxide ion. NaOH is an ionic base and a strong electrolyte; therefore, it will dissociate completely. This means the $[NaOH] = [OH^-]$. To calculate the concentration, we need to convert the mass of NaOH to moles (NaOH = 40.0 g/mol).

$$\text{moles} = 0.40 \text{ g}\left(\frac{1 \text{ mol}}{40.0 \text{ g}}\right) = 0.010 \text{ mol NaOH}$$

$$\left[NaOH\right] = \left[OH^-\right] = 0.010 \text{ mol/L}$$

Step 2: Substitute into the equilibrium law and solve for hydronium ion concentration.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$K_W = [H_3O^+][OH^-]$$

$$\left[H_3O^+\right] = \frac{K_W}{\left[OH^-\right]} = \frac{1.0 \text{ x } 10^{-14}}{1.0 \text{ x } 10^{-4} \text{ mol/L}} = 1.0 \text{ x } 10^{-10} \text{ mol/L}$$

Learning Activity 5.3: K_W Calculations

1. If the $[OH^-]$ in a sodium hydroxide solution is 0.050 mol/L, what is $[H_3O^+]$?
2. 0.250 mole of hydrogen chloride gas is dissolved in 2.00 L of water. Write the dissociation equation for this gas and calculate both $[H_3O^+]$ and $[OH^-]$.
3. 10.0 g of lithium hydroxide is dissolved in 750.0 mL of water. Write the dissociation equation and calculate both $[H_3O^+]$ and $[OH^-]$.
4. 10.0 g of calcium hydroxide is dissolved in 400.0 mL of solution. Write the dissociation equation and calculate both $[H_3O^+]$ and $[OH^-]$.
5. If the $[H_3O^+]$ of a barium hydroxide solution is $1.00 \text{ x } 10^{-13}$ mol/L, calculate the $[OH^-]$. How many grams of barium hydroxide are needed to make a litre of this solution?
6. Calculate the $[H_3O^+]$ in milk of magnesia (magnesium hydroxide) that has an $[OH^-]$ of $1.43 \text{ x } 10^{-4}$ mol/L.

Check the answer key.

Lesson Summary

In this lesson, you learned that, at 25°C, the value of the ion-product constant for water, K_W, is $1.0 \text{ x } 10^{-14}$. You used this constant to solve different K_W problems and to determine the concentration of hydroxide ions and hydronium ions. In the next lesson, you will use Le Châtelier’s principle to describe how an acid-base indicator works in terms of colour shifts.

Assignment 5.3: Solving K_W Problems (10 marks)

Solve each of the following K_W problems. For each solution, set up the correct equation and solve for the unknown. Don't forget to include the correct units and significant figures for your final answer. *(1 mark for setting up the correct equation, 1 mark for substituting variables, 1 mark for the answer with correct units and significant figures)*

1. What is the hydroxide ion concentration in a solution with a hydronium concentration of 6.80×10^{-10} mol/L? *(3 marks)*

2. What is the $[H_3O^+]$ in a solution with $[OH^-]$ of 5.67×10^{-3} mol/L? *(3 marks)*

continued

Assignment 5.3: Solving K_W Problems (continued)

3. If the nitric acid (HNO_3) concentration is 0.0020 mol/L, what is the $[OH^-]$? *(4 marks)*

LESSON 4: LE CHÂTELIER'S PRINCIPLE (3 HOURS)

Lesson Focus

SLO C12-5-04: Perform a laboratory activity to formulate an operational definition of pH.

SLO C12-5-05: Describe how an acid-base indicator works in terms of colour shifts and Le Châtelier's principle.

SLO C12-5-06: Solve problems involving pH.

Lesson Introduction

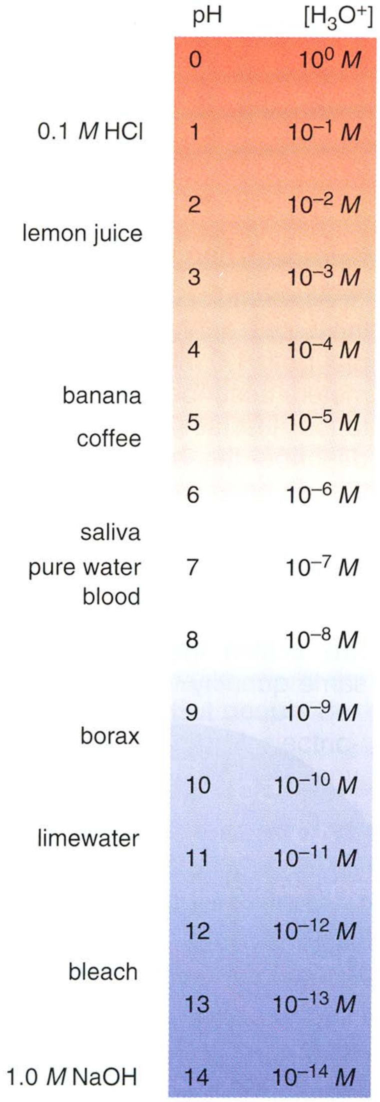
Have you seen a commercial advertising a shampoo that is pH balanced? Some flowers will bloom different colours, based upon the soil's pH. Swimming pools and aquariums require the water's pH to be kept in a very narrow range. In this lesson, you will learn what pH really means and how it can be measured.

The pH Scale

You may have been introduced to the pH scale in a previous science course. If so, you know that it is a numeric scale which ranges from 0 to 14. Substances tending to the 0 end of the scale are increasingly acidic, while those tending to the 14 end of the scale are increasingly basic. Neutral solutions have a pH of 7. The usual range of pH values encountered is between 0 and 14, with 0 being the value for concentrated hydrochloric acid (1 mol/L HCl), 7 being the value for pure water (neutral pH), and 14 being the value for concentrated sodium hydroxide (1 mol/L NaOH).

In theory, any pH is possible, especially at elevated temperatures where pure water decreases in pH because of increased ionization (but it is still neutral, neither acidic nor alkaline). In practice, there are limitations because not everything will dissolve at sufficiently high concentration and ionize sufficiently at that concentration to achieve any given pH. It is possible to get a pH of -1 with 10 mol/L HCl, but that is the practical limit of acidity. At the other extreme, a 10 mol/L solution of NaOH would have a pH of 15.

Remember that you can view this and other graphics in colour in the learning management system (LMS).



LeMay, H. Eugene, Jr., *et al. Chemistry: Connections to Our Changing World.* 2nd ed. Upper Saddle River, NJ: Prentice-Hall Inc., 2000. 629.

Every water solution is neutral, acidic, or basic.

- A **neutral** solution occurs when the hydronium ion concentration is **equal** to the hydroxide ion concentration. $[H_3O^+] = [OH^-]$
- An **acidic** solution occurs when the hydronium ion concentration is **greater** than the hydroxide ion concentration. $[H_3O^+] > [OH^-]$
- A **basic** solution occurs when the hydronium ion concentration is **less** than the hydroxide ion concentration. $[H_3O^+] < [OH^-]$

Working with hydrogen-ion concentrations in molarity can be cumbersome, as most concentrations are very small (i.e., 4×10^{-8} mol/L or 0.00000004 mol/L). For this reason, Soren P. Sorenson of Denmark proposed an easier method, called the pH scale. Actual pH and concentration is calculated by using some simple rules.

Logarithms

In case you are unfamiliar with logarithms, this tutorial will help you understand the rest of this lesson and complete the learning activity and assignment.

As with exponents, **logarithms** (or **logs**) are a way of working with very large and very small numbers. A number's base 10 or common logarithm is its exponent when 10 is the base. For example, for the number 10 000,

$$\log(10\ 000) = 4 \qquad 10\ 000 = 10^{4}$$

The logs of other numbers are not whole numbers, but the process is the same.

$$\log(250) = 2.3979 \qquad 250 = 10^{2.3979}$$

A number greater than one represents a positive log and a number less than one results in a negative log.

$$\log(0.001) = -3 \text{ or } 0.001 = 10^{-3}$$

$$\log(0.000\ 25) = -3.602 \text{ or } 0.000\ 25 = 10^{-3.602}$$

To calculate "log(0.000 25)" on your calculator, one of these should be the sequence of steps:

- press **LOG** then enter 0.000 25
- press =

or

- enter 0.000 25
- press **LOG**

Just to make sure you have the hang of it, find the log of each of the following values:

a. $1.3 \text{ x } 10^{-5}$

b. $7.2 \text{ x } 10^{-11}$

c. 0.0054

Answers

a. $\log(1.3 \text{ x } 10^{-5}) = -4.89$

b. $\log(7.2 \text{ x } 10^{-11}) = -10.14$

$\log(0.0054) = -2.27$

pH Calculations

For this lesson, assume that all acids and bases completely ionize or dissociate. This means that the hydronium ion concentration is equal to the concentration of a given monoprotic acid. It also means the hydroxide ion concentration is equal to the concentration of a given monohydroxic base.

Animations of various acidic and basic solutions, and of reading their pH values on the pH metre, can be viewed at www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animationsindex.htm.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

pH

$\text{pH} = -\log [H_3O^+]$ (all in base 10) or, $\text{pH} = -\log [H^+]$

Note

The pH of a solution is the negative logarithm of the hydrogen ion (hydronium ion) concentration.

In a neutral solution, the $[H^+] = 1 \text{ x } 10^{-7}$ mol/L. The pH of a neutral solution is

$$\text{pH} = -(\log \text{H+}) = (-\log 10^{-7}) = 7.0$$

Since the pH is an integer number, it is also possible to directly write the value of $[H^+]$. If the hydrogen ion (or hydronium ion) concentration is written in scientific notation and has a coefficient of 1, then the pH of the solution equals the exponent.

Example 1

If $[H^+] = 1 \times 10^{-13}$ mol/L, then the corresponding pH is 13.0

Example 2

Calculate the pH of an HCl solution with a concentration of 5.0×10^{-6} mol/L.

Step 1: Calculate the log of the $[H_3O+]$.

$$\text{pH} = -\log[H_3O^+]$$

$$\text{pH} = -\log(5.0 \times 10^{-6})$$

Step 2: Change the sign to get the negative log.

$$\text{pH} = -(-5.30) = 5.3$$

Round the log to two significant digits.

Example 3

The pH of a solution is 3.25. Calculate the hydronium ion concentration in the solution. For this solution, you need to use the opposite of the log, called the **antilog** or **inverse log**. On your calculator, the antilog function is usually above the log, **10^x**.

$$[H_3O^+] = 10^{-\text{pH}}$$

When calculating the antilog, be sure to use the "+/–" button to make the negative of the pH value. On most calculators; enter 3.25, press +/–, and press the 10^x button.

$$[H_3O^+] = 10^{-3.25}$$

$$[H_3O^+] = 5.62 \times 10^{-4} \text{ mol/L}$$

Again, round the concentrations to three significant digits, in this case.

pOH

The pOH of a solution equals the negative logarithm of the **hydroxide ion** concentration. In a neutral solution, the $[OH^-] = 1 \times 10^{-7}$ mol/L. The pOH of a neutral solution is

$$pOH = -\log [OH^-] \text{ (all in base 10)}$$

$$pOH = -(\log OH^-) = (-\log 10^{-7}) = 7.0$$

Note

The pOH of a solution is the negative logarithm of the hydroxide ion concentration.

Since the hydronium and hydroxide ions are in equilibrium, and $K_W = 1 \times 10^{-14}$,

$$pH + pOH = 14$$

This simple equation can be manipulated to solve for unknown pH or pOH as follows:

$$pH = 14 - pOH$$

$$pOH = 14 - pH$$

Example 4:

The pH of a solution is 10.30. What is the hydroxide ion concentration?

Step 1: Find the pOH

$$pOH = 14.00 - pH$$

$$pOH = 14.00 - 10.30 = 3.70$$

Step 2: Convert pOH to hydroxide ion concentration using the antilog function. Make this conversion with pOH just as you would with the pH.

$$[OH^-] = 10^{-pOH}$$

$$[OH^-] = 10^{-3.70}$$

$$[OH^-] = 2.0 \times 10^{-4} \text{ mol/L}$$

Example 5

What is the pH of a 5.0×10^{-5} mol/L $Mg(OH)_2$ solution?

To calculate the pH, you need the hydronium ion concentration. You can calculate hydronium from the hydroxide ion concentration.

Step 1: Calculate the hydroxide ion concentration. Assume the $Mg(OH)_2$ dissociates completely.

$$Mg(OH)_{2(s)} \rightarrow Mg^{2+}{}_{(s)} + 2\,OH^{-}{}_{(s)}$$

$$[OH^-] = 2 \times [Mg(OH)_2] = 2 \times 5.0 \times 10^{-5}\text{ mol/L} = 1.0 \times 10^{-4}\text{ mol/L}$$

Step 2: You can use the ion product of water to calculate the hydronium ion concentration.

$$K_W = [H_3O^+][OH^-]$$

$$\left[H_3O^+\right] = \frac{K_W}{\left[OH^-\right]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}\text{ mol/L}} = 1.0 \times 10^{-10}\text{ mol/L}$$

Step 3: Calculate pH.

$$pH = -\log[H_3O^+] = -\log(1.0 \times 10^{-10}) = 10.00$$

An alternative method would be to determine the pOH and then subtract from 14 to determine the pH.

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-4}) = 4.00$$

$$pH = 14.00 - 4.00 = 10.00$$

Indicators

Indicators are weak, organic acids that change colour when the hydronium or hydroxide ion concentration is changed. Indicators change colour over a given pH range. (The colour of each indicator can be coordinated with the pH.) The colour of the indicator can be compared to a standard to determine the pH of the solution.

Often, an indicator (or a combination of several indicators) can be incorporated into paper or plastic strips, making their use more convenient. You just dip the strip into the solution and the strip changes colour, depending on the indicators used and the pH of the solution. Note that using an indicator is not as accurate as a pH meter. If you have a pool or hot tub at home, you probably use a kit that contains an acid-base indicator to determine the water's pH. On the following page is a table of some common indicators and the pH range they measure.

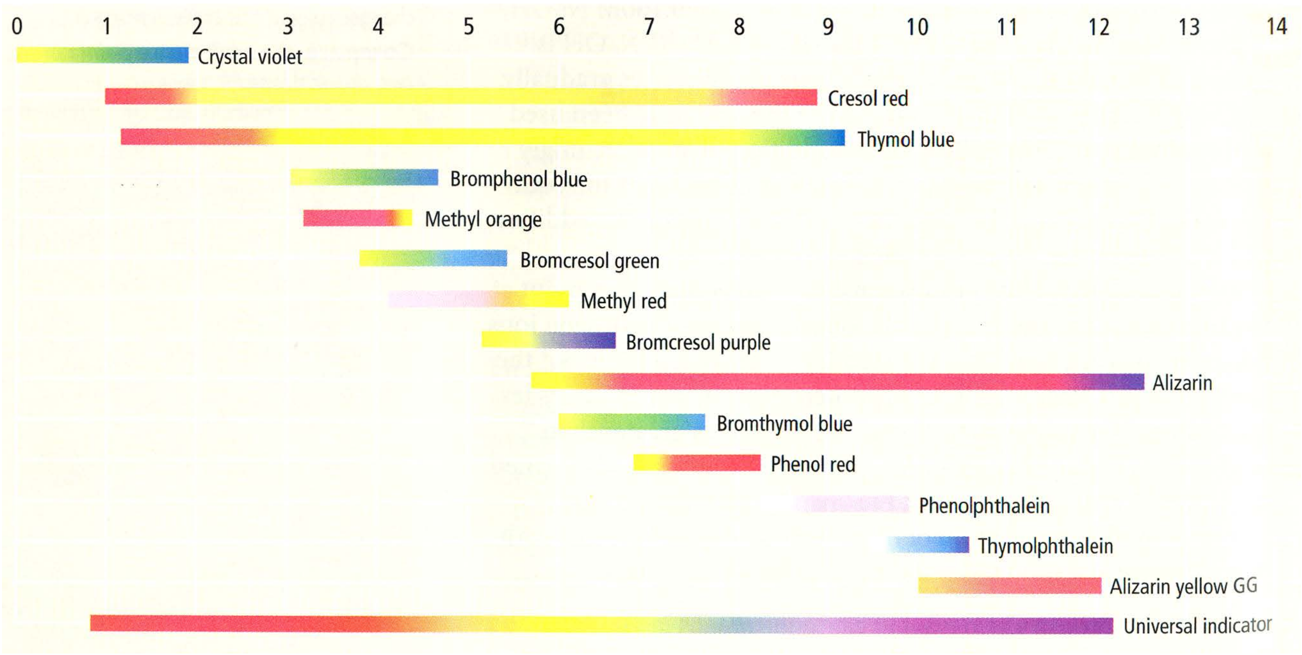


Photo Credit: Matt Meadows.
Buthelezi, Thandi, *et al. Chemistry: Matter and Change: Student Text.* New York, USA: The McGraw-Hill Companies, Inc., Glencoe Science, 2008. 662.

Le Châtelier's Principle Revisited

You learned in the last module that changing the reactant or product concentrations disturbs equilibrium in the following ways.

- **Adding a reactant** always shifts a reversible reaction in the direction of **products**.
- **Removing a reactant** always shifts a reversible reaction in the direction of **reactants**.
- **Adding a product** causes the reaction to shift in the direction of the formation of the **reactants**.
- **Removing a product** shifts the reaction in the direction of the formation of the **products**.

Le Châtelier's principle can be used to explain colour changes, also called colour shifts.

Colour 1 Colour 2
↓ ↓

$$\mathrm{HIn}_{(aq)} \rightleftharpoons \mathrm{H}^{+}_{(aq)} + \mathrm{In}^{-}_{(aq)}$$

(acid form) (basic form)

The presence of an acid increases the hydrogen ion concentration, which causes a shift from colour 2 toward colour 1. The presence of a base decreases the hydrogen ion concentration, causing a shift from colour 1 toward colour 2.

The range of change is often about 2 pH units, although quite a few are less. The human eye responds more readily to some shades of colour than to others, and some substances are naturally more intensely coloured than others, even at the same concentration.

The Universal Indicator

Another type of indicator that can be used in aqueous solutions of acids and bases is the universal indicator. This indicator will change through a series of colours at specific pH ranges. This can be quite useful if the pH of a solution is unknown and is of interest. The indicator can be added to any solution and will turn to one of the colours below. The chart below is used to determine the relative pH of any solution.

pH Range	Description	Colour
0-3	Highly Acidic	Red
3-6	Acidic	Orange/Yellow
7	Neutral	Green
8-11	Basic	Blue
11-14	Highly Basic	Purple

The Natural Universal Indicator: Red Cabbage Juice

The universal indicator works quite nicely to determine the relative pH of any solution. An alternate to the universal indicator is red cabbage juice. Red cabbage juice can be used to achieve similar results. When finely chopped and boiled in water, red cabbage will produce a solution that is slightly acidic. This acid will behave as an indicator, which will produce differently coloured solutions at different pH ranges. The chart below summarizes these pH ranges and the expected colours.

pH Range	Description	Colour
0-3	Highly Acidic	Red
3-6	Acidic	Purple
6-8	Neutral	Bluish Purple
8-11	Basic	Bluish Green
11-14	Highly Basic	Greenish Yellow

By following the instructions in the following home-based laboratory activity, you can test some items that you might have around your home to determine the relative pH of these various solutions. If you do not have access to these items, use the following link to help answer the questions in the learning activity: www.dlt.ncssm.edu/core/Chapter16-Acid-Base_Equilibria/red_cabbage.htm.

You can also find this video on the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chemvid7.htm. It is titled ***Red Cabbage Indicator***.

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Learning Activity 5.4: The pH Scale Laboratory Activity

Materials

- red cabbage leaves
- water
- any six household solutions to test, for example:
  - vinegar
  - laundry detergent
  - lemon juice
  - bleach
  - drain cleaner
  - glass cleaner (e.g., Windex)
- a pot in which to boil water
- a strainer

Procedure

1. Take half of a red cabbage and chop it into finely divided pieces. Place the chopped cabbage and 4 cups of water in a pot and bring to a boil. Boil for 3 to 5 minutes, or until the leaves become quite limp and the water has turned to a bright red colour.
2. Strain the cabbage leaves and collect the solution. Discard the cabbage into a composter if you have one. Pour the solution into 6 to 8 small containers.

continued

Learning Activity 5.4: The pH Scale Laboratory Activity (continued)

3. Add a small amount of one of your six household solutions to one container. Repeat this step for each of your household solutions. Record your results.
4. Use the results you collected (or the results from www.dlt.ncssm.edu/core/Chapter16-Acid-Base_Equilibria/red_cabbage.htm) to answer the questions in the learning activity.

You can also find this video on the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chemvid7.htm. It is titled ***Red Cabbage Indicator***.

Part A: Calculations and Short Answer

1. Determine the pH of each of the following values
   a. $[H^+] = 2.0 \times 10^{-3}$
   b. $[H^+] = 3.5 \times 10^{-4}$
   c. $[OH^-] = 7.5 \times 10^{-7}$
   d. $[OH^-] = 9.0 \times 10^{-9}$
2. Determine the concentration of $[H^+]$ for each of the following.
   a. pH = 7.0
   b. pOH = 5.0
   c. pH = 13.0
   d. pOH = 11.0
3. Determine the pOH of the following if the pH is given.
   a. pH = 5.4
   b. pH = 8.6
4. Determine the concentration of the OH^- ions in the solutions with the following pH values.
   a. pH = 1.6
   b. pH = 9.2
5. Calculate the pH of each of the following solutions.
   a. 0.0020 mol/L HCl
   b. 4.00×10^{-4} mol/L $Ba(OH)_2$

continued

Learning Activity 5.4: The pH Scale Laboratory Activity (continued)

6. Which indicator changes colour in a solution that has a pH of 2? Use the table provided in the lesson to help you answer this question.
7. Which indicator(s) would you choose to show that a solution has changed from a pH of 3 to a pH of 5? From a pH 4.6 to a pH of 4.9?
8. Red cabbage juice can be used as an indicator to show the relative pH ranges of different solutions. Use the data that you collected previously as well as your knowledge of acids and bases to answer each of the following questions:
   a. CLR is a commonly used household cleaner. A student tests CLR in a solution of red cabbage juice and the solution turns a bright red colour. What would be the relative pH range of CLR?
   b. A clear, carbonated beverage is tested in red cabbage juice and the solution turns a light purple colour. What would be the pH range of the beverage?
   c. An antacid tablet is added to red cabbage juice and the solution turns a light green colour. What would be the pH range of the antacid?
9. Are there any OH^- ions in an acidic solution with a pH of 3.25? How do you know?

Part B: Multiple Choice

1. Which of these solutions is the **most** basic?
   a. $[H^+] = 1 \times 10^{-2}$
   b. $[OH^-] = 1 \times 10^{-4}$
   c. $[H^+] = 1 \times 10^{-11}$
   d. $[OH^-] = 1 \times 10^{-13}$
2. In a neutral solution, the $[H^+]$ is
   a. 10^{-14}
   b. 0
   c. 1×10^{7} mol/L
   d. Equal to $[OH^-]$

continued

Learning Activity 5.4: The pH Scale Laboratory Activity (continued)

3. A solution in which the hydroxide ion concentration is 1×10^{-8} mol/L is
   a. acidic
   b. basic
   c. neutral
4. In an acidic solution,
   a. $[H^+]$ is equal to $[OH^-]$
   b. $[H^+]$ is greater than $[OH^-]$
   c. $[H^+]$ is less than $[OH^-]$
   d. $[H^+]$ are both $[OH^-] = 1 \times 10^{-7}$ mol/L

Check the answer key.

Lesson Summary

In this lesson, you learned how an acid-base indicator works in terms of colour shifts and Le Châtelier’s principle. You also practiced solving several problem types involving pH. In the next lesson, you will distinguish between strong and weak acids and bases.

Notes

Assignment 5.4: Solving pH Problems (23 marks)

1. Determine the pH of each of the following values. *(8 marks)*
   a. $[H^+] = 1.0 \times 10^{-6}$ *(1 mark)*

   b. $[H^+] = 1.0 \times 10^{-8}$ *(1 mark)*

   c. $[OH^-] = 1.0 \times 10^{-2}$ *(3 marks)*

   d. $[OH^-] = 1.0 \times 10^{-5}$ *(3 marks)*

2. Determine the concentration of $[H^+]$ of each of the following. *(4 marks)*
   a. $pH = 3.4$ *(2 marks)*

continued

b. pH = 8.9 *(2 marks)*

3. Determine the concentration of the OH^- ions in the solutions with the following pH values. *(4 marks)*

a. pH = 4.1 *(2 marks)*

b. pH = 5.1 *(2 marks)*

4. Calculate the pH for each of the following solutions. *(5 marks)*

a. 0.050 mol/L NaOH *(3 marks)*

b. 0.50 mol/L HNO_3 *(2 marks)*

continued

Assignment 5.4: Solving pH Problems (continued)

5. Explain why it is useful to use a universal indicator or cabbage juice with common household items. *(2 marks)*

Notes

LESSON 5: STRONG AND WEAK ACIDS AND BASES (2 HOURS)

Lesson Focus

SLO C12-5-07: Distinguish between strong and weak acids and bases.
Include: electrolytes and non-electrolytes

SLO C12-5-08: Write the equilibrium expression (K_a or K_b) from a balanced chemical equation.

Lesson Introduction

Up to this point, you have learned both theoretical and mathematical aspects of equilibrium. In this lesson, you will learn to distinguish between strong and weak acids and bases. You will also practice writing equilibrium expressions from a balanced chemical equation.

Strong versus Concentrated

In Grade 11 Chemistry, you learned the difference between dilute and concentrated. A **dilute solution** contains a small amount of solute per volume (for example, 0.0010 mol/L) and a concentrated solution contains a large amount of solute per volume (such as 11.2 mol/L). Are strong and concentrated solutions synonymous? What about weak and dilute solutions? Obviously, a dilute solution of a strong acid is possible (0.0010 mol/L of sulphuric acid) as is a concentrated solution of a weak acid (17.4 mol/L of acetic acid). Acids and bases differ greatly in their strength.

Strong Acids

In general, a **strong acid** will dissociate close to 100% and have a very large K_{eq}. This means that the reaction goes to completion with very little, if any, of the reactant HA left.

$$HA + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + A^-_{(aq)}$$

A single arrow is used.

Chemists do not usually write equilibrium expressions for strong acids and bases because there is essentially no equilibrium. The acid-dissociation equilibrium constant (K_a) is the mathematical product of the equilibrium concentrations of the products of this reaction divided by the equilibrium concentration of the original acid.

Note

For the purposes of this course, you can use K_a or K_{eq} when writing these equilibrium expressions.

And so, the K_a is the product of the concentrations of the products over the concentration of the reactants (not including water).

$$K_a = \frac{[H_30^+][A^-]}{[HA]}$$

At equilibrium, the K_{eq} for any strong acid is very large, (K_{eq} >>>> 1). In the example of a strong acid, such as hydrochloric acid, there are virtually no HCl molecules present in the aqueous solution of acid.

$$\underbrace{HCl_{(aq)}}_{\substack{\text{0.0004\% at} \\ \text{equilibrium}}} + H_2O_{(l)} \rightarrow \underbrace{H_3O^+_{(aq)} + Cl^-_{(aq)}}_{\substack{\text{99.996\% at} \\ \text{equilibrium}}}$$

Other examples of strong acids are

- $HClO_4$ (perchloric acid)
- HI (hydroiodic acid)
- HBr (hydrobromic acid)
- HNO_3 (nitric acid)
- H_2SO_4 (sulphuric acid)

Strong Bases

A **strong base** is one that also completely dissociates into ions. Examples of strong bases are

- NaOH (sodium hydroxide)
- KOH (potassium hydroxide)
- LiOH (lithium hydroxide)
- $Ca(OH)_2$ (calcium hydroxide)
- RbOH (rubidium hydroxide)
- $Ba(OH)_2$ (barium hydroxide)

Note

In both strong acids and strong bases, the reaction is so far to the right that there is essentially no reactant left and so there is no equilibrium.

For strong acids and bases, the reactions use only a forward arrow since there is no reverse reaction. For example,

0.50 mol/L of HCl will produce $[H^+] = [Cl^-] = 0.50$ mol/L

0.50 mol/L of NaOH will produce $[Na^+] = [OH^-] = 0.50$ mol/L

Weak Acids

Weak acids dissociate only slightly into ions; therefore, these dissociation equations use a reversible arrow, unlike reactions involving strong acids and bases.

$$\underbrace{HC_2H_3O_{2(aq)}}_{\text{98.7\% at equilibrium}} + H_2O_{(l)} \rightleftharpoons \underbrace{H_3O^+{}_{(aq)} + C_2H_3O_2^-{}_{(aq)}}_{\text{1.3\% at equilibrium}}$$

In the above reaction, very little product is formed, meaning the reverse reaction is preferred, and K_{eq} is very small.

$K_{eq} <<<< 1$

The equilibrium expression can be simplified since the concentration of water is very large compared to the concentration of the acid. The equilibrium expression is

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons A^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

> **Note**
>
> The value of K_a or K_b provides a measure of the relative strength of an acid or a base.

Weak Bases

Like their weak acid counterparts, **weak bases** dissociate only slightly into ions. An important weak base is ammonia. Note that when writing the dissociation of a weak base, you must again use the reversible arrow.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Using the reaction above, you can now write the equilibrium expression.

$$K_{eq} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Similar to a weak acid, this equilibrium expression can be simplified, since the concentration of water is very large compared to the concentration of the base. As a result, we can write the following:

$$K_{eq} = K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \text{ x } 10^{-5}$$

We use the symbol K_b for the equilibrium constant for a base.

Other examples of weak bases are

- $C_6H_5NH_2$ (aniline)
- CH_3NH_2 (methylamine)
- C_5H_5N (pyridine)

Note

Stronger acid = higher % dissociation = higher [H_3O^+] = larger K_a

Stronger base = higher % dissociation = higher [OH^-] = larger K_b

Weaker acid = lower % dissociation = lower [H_3O^+] = smaller K_a

Weaker base = lower % dissociation = lower [OH^-] = smaller K_b

Example 1

For the following reaction, place a single arrow (→) between the reactants and products if the reaction goes to completion. Place a double-sided arrow ($\rightleftharpoons$) if the reaction reaches equilibrium. Then, write the equilibrium expression.

$$HNO_{3(aq)} + H_2O_{(l)} \qquad NO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

Solution: Since nitric acid is a strong acid, the reaction will go to completion.

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow NO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

$$K_{eq} = \frac{\left[NO_3{}^-\right]\left[H_3O^+\right]}{\left[HNO_3\right]}$$

In Appendix D at the end of this course, there is a K_a chart for acids, called the **Relative Strengths of Acids**. The larger the K_a, the stronger the acid and the greater the tendency to release H^+ (H_3O^+) ions into solution. Using this argument, the species on the other side of the arrow are bases. They have a tendency to pick up H^+ (H_3O^+). If the strongest acids are on the top left, then the strongest bases must be toward the bottom right. Looking at the table, the amide ion ($NH_2{}^-$) is found at the bottom of the table on the right-hand side; for this reason, it is the strongest base species.

$$NH_3 + H_2O \rightarrow H_3O^+ + NH_2{}^-$$

This is closely followed by the oxide ion (O^{2-}), which is found just above the reaction involving ammonia.

$$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$$

Example 2

Use the K_a table (Appendix D) to answer the following questions regarding the reaction between solutions of HSO_4^- and $H_2PO_4^-$.

1. Using the K_a table, which of the above species will behave as an acid when combined together?
2. Write a balanced chemical equation between the two ions. Be sure to include state symbols.
3. Identify the Brønste-Lowry acid/base, conjugate acid/base pairs.
4. Determine which acid is the strongest and, therefore, whether the forward or reverse reaction would be favoured.

Solution:

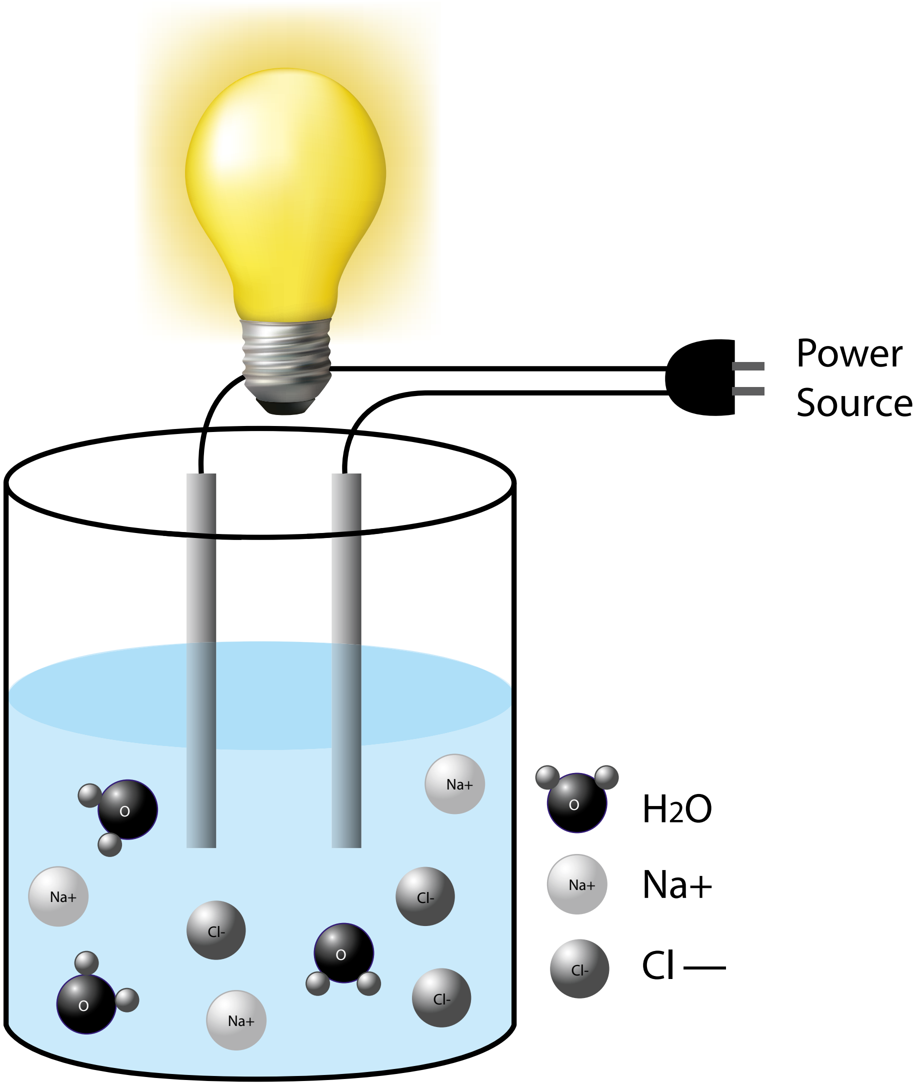
1. Of the two species, HSO_4^- would behave as an acid and is more likely to donate the hydrogen ion in solution. HSO_4^- has a larger K_a value, which means that it dissociates into its ions to a greater extent than $H_2PO_4^-$.
2. $HSO_4^-{}_{(aq)} + H_2PO_4^-{}_{(aq)} \rightleftharpoons H_3PO_{4(aq)} + SO_4^{2-}{}_{(aq)}$. Note that a double arrow is used to signify that the solution will be in a state of equilibrium and that completion of the reaction will not be achieved.
3. Because HSO_4^- donates the H^+ ion into solution, it is the acid. $H_2PO_4^-$ is the hydrogen ion acceptor and is, therefore, behaving as a base in this reaction. The conjugate base of HSO_4^- would be SO_4^{2-} and the conjugate acid of $H_2PO_4^-$ is H_3PO_4.
4. Using your table, compare H_3PO_4 and HSO_4^-. Because HSO_4^- has a larger K_a value and is a stronger acid, the forward reaction is favoured. The stronger acid will always determine the direction of the reaction as it will ionize to a greater extent.

Weak and Strong Electrolytes

The next time you see an advertisement for a sports drink, pay attention to how they use the term "electrolyte replacement." An **electrolyte** is defined as a compound that, in its molten or aqueous state, conducts an electric current. In your body, nervous signals are electrical in nature and require electrolytes to be transmitted.

All ionic compounds, like acids and bases, are electrolytes since they will dissociate 100% in water. Electrolytes are classified as strong or weak. A **strong electrolyte** is one that produces many ions in solution. Sodium chloride is an example of a strong electrolyte.

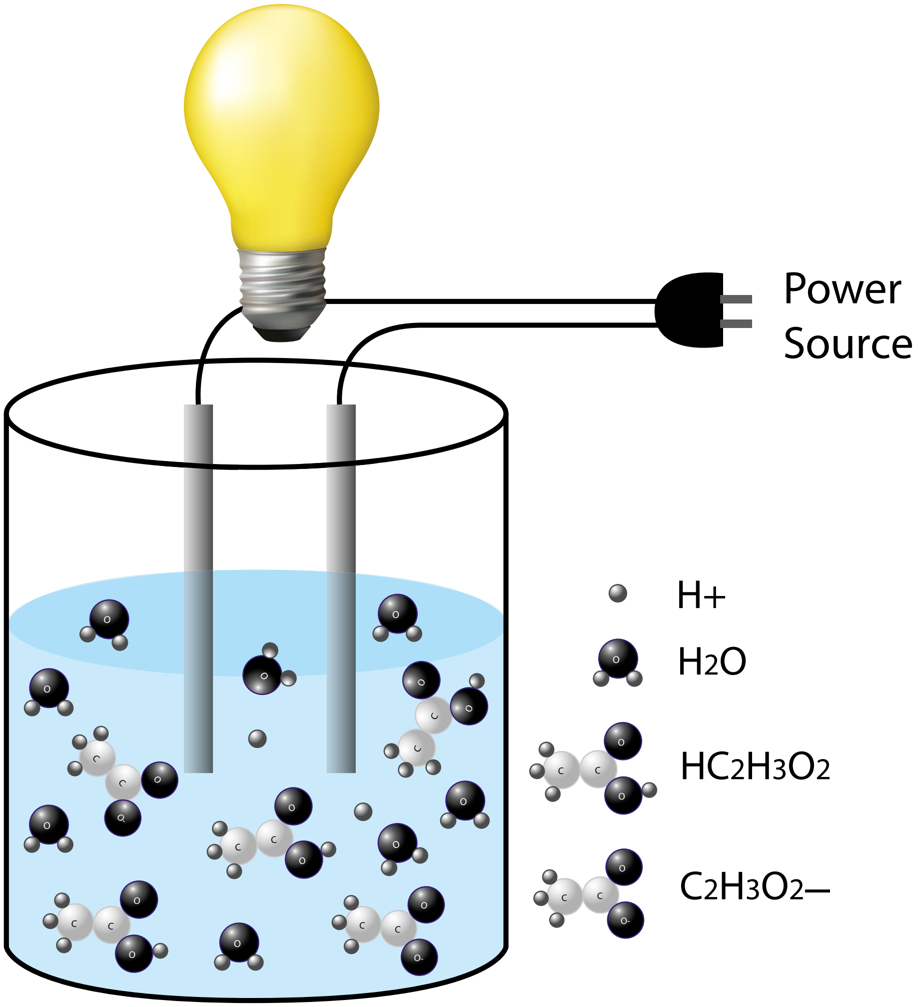
Sodium Chloride in Water



A **weak electrolyte** will conduct only a small electric current because it does not produce many ions in solution. A weak electrolyte will make the light bulb on our conductivity tester glow very dimly. Weak electrolytes do not produce many ions because they only partially dissociate. These are usually polar covalent molecules that partially ionize. An example of a weak electrolyte is acetic acid, commonly called vinegar (when diluted). The reaction that follows was also presented in the previous section.

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightarrow C_2H_3O_2{}^{-}{}_{(aq)} + H_3O^{+}{}_{(aq)}$$

Acetic Acid in Water



Learning Activity 5.5: Strong and Weak Acids and Bases

Use the Relative Strengths of Acids Table (Appendix D) found in the appendices section at the back of this course to complete these questions.

1. For each of the following:
   - Identify which of the two would behave as the acid and which would behave as the base.
   - Write balanced chemical equations for the reactions.
   - Indicate the acid/base conjugate acid base pairs.
   - Determine which direction would be favoured in each, using the K_a table.

   a. $H_3PO_4 + C_2H_3O_2^-$
   b. $C_2H_3O_2^- + HSO_4^-$
   c. $SO_3^{2-} + NH_4^+$
   d. $HO_2^- + H_2CO_3$

2. Using the following ions:
   - CO_3^{2-}, HCO_3^-, HPO_4^{2-}, CH_3COO^-, HSO_3^-

   a. Identify which are acids and list in order of decreasing (strongest to weakest) strength.
   b. List those that behave as a base in order of decreasing strength.

3. List the following substances in increasing order of strength (weakest to strongest) as acids and bases. Some may be used more than once.
   - F^-, $H_2PO_4^-$, HSO_4^-, H_2S, $C_2H_3O_2^-$, HNO_2, HTe^-
     — Acids
     — Bases

Check the answer key.

Lesson Summary

In this lesson, you learned the difference between strong and weak acids and bases. In doing so, you were introduced to substances that are electrolytes and those that are non-electrolytes. Finally, you saw how to write equilibrium expressions from a balanced chemical equation.

Assignment 5.5: Acid and Base Equilibrium (10 marks)

1. Why is spilling 1 mol/L acetic acid relatively harmless compared to spilling 1 mol/L sulphuric acid? *(2 marks)*

2. Are all strong acids also strong electrolytes? Explain your answer. *(2 marks)*

continued

Assignment 5.5: Acid and Base Equilibrium (continued)

3. For the following reactions:

- Place a single arrow (→) between the reactants and products if the reaction goes to completion. Place a double-sided arrow ($\rightleftharpoons$) if the reaction reaches equilibrium. *(1 mark each)*
- Write the equilibrium expression. *(2 marks each)*

a. $H_2O + HC_2H_3O_2 \qquad\qquad H_3O^+ + C_2H_3O_2^-$

b. $NH_3 + H_2O \qquad\qquad NH_4^+ + OH^-$

LESSON 6: SOLVING K_a AND K_b PROBLEMS (4 HOURS)

Lesson Focus

SLO C12-5-09: Use K_a or K_b to solve problems for pH, percent dissociation, and concentration.

Lesson Introduction

In the previous lesson, you learned how to write equilibrium expressions for acid-base reactions. In this lesson, you will use the equilibrium expression problem to solve for pH, percent dissociation, and concentration.

Weak Acid/Weak Base Problems

There are several different problems that you might encounter. Basically, there are two types of questions for a weak acid/weak base:

Calculating the Dissociation Constant

Given the initial concentration of the acid/base and percent dissociation, pH, pOH, $[H_3O^+]$ or $[OH^-]$, find K_a or K_b.

Example 1

Using the equilibrium concentration 0.75 mol/L solution of a weak base (NH_3) and $[OH^-] = 1.0 \times 10^{-4}$ mol/L, find K_b.

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)}$$

Step 1: Write the equilibrium expression.

$$K_b = \frac{[NH^+_4][OH^-]}{[NH_3]}$$

Step 2: Substitute the given values.

$[NH^+_4] = [OH^-] = 1.0 \times 10^{-4}$ mol/L since the stoichiometry is 1:1.

$$K_b = \frac{\left[1.0 \times 10^{-4}\right]\left[1.0 \times 10^{-4}\right]}{[0.75]}$$

$$K_b = 1.3 \times 10^{-8}$$

Example 2

A 0.10 mol/L solution of acetic acid is only partly ionized. If, at equilibrium, the hydronium ion concentration is 1.3×10^{-3} mol/L, what is the acid dissociation constant, K_a?

In this problem, you are given the initial acid concentration and the equilibrium concentration of the hydronium ions. You need the equilibrium concentration of the acetic acid and the acetate ions. As in other equilibrium problems, use an ICE table to determine equilibrium concentrations.

Step 1: Write out the dissociation equation.

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

Step 2: Determine the concentration of ions. Since the acetic acid dissociates and the molar ratio of hydronium to acetate is 1:1,

$[H_3O^+] = [C_2H_3O_2^-] = 1.3 \times 10^{-3}$ mol/L.

Step 3: Set up an ICE table. The acetic acid concentration will be reduced by the amount of hydronium and acetate that are produced.

	$HC_2H_3O_2$	$+ H_2O$	$\rightleftharpoons$	H_3O^+	$C_2H_3O_2^-$
I	0.10			0.0	0.0
C	$- 1.3 \times 10^{-3}$			$+ 1.3 \times 10^{-3}$	$+ 1.3 \times 10^{-3}$
E	$0.987 \approx 0.99$			1.3×10^{-3}	1.3×10^{-3}

Note that 0.987 is rounded to 0.99 because of significant figures.

Step 4: Substitute into equilibrium law.

$$K_b = \frac{\left[H_3O^+\right]\left[C_2H_3O_2^-\right]}{\left[HC_2H_3O_2\right]}$$

$$K_b = \frac{\left[1.3 \text{ x } 10^{-3}\right]\left[1.3 \text{ x } 10^{-3}\right]}{\left[0.099\right]} = 1.7 \text{ x } 10^{-5}$$

Calculating the Concentration of Dissociated Species

Given the initial concentration of the acid/base and K_a or K_b, find $[H_3O^+]$, $[OH^-]$, percent dissociation, and pH.

This type of problem requires you to know the acid or base dissociation constant and then calculate all the species in the solution. This requires the use of the "ICE" table.

Example 3

HA is a weak acid with a K_a of 7.3 × 10^{-8}. What is the concentration of all species (HA, H_3O^+, and A^-) if the initial concentration of HA is 0.50 mol/L?

Step 1: Write out the dissociation equation and plot an ICE table. Since you do not know how much of the acid ionizes, you do not know the equilibrium concentrations. As you did in the last module, use the variable, *x*, to represent the amount that ionizes.

	HA	+ H_2O	$\rightleftharpoons$	H_3O^+	A^-
I	0.50			0.0	0.0
C	–x			+ x	+ x
E	0.50–x			x	x

Step 2: Write the equilibrium law and substitute the values from the ICE table.

$$K_a = \frac{\left[H_3O^+\right]\left[A^-\right]}{\left[HA\right]}$$

$$7.3 \text{ x } 10^{-8} = \frac{[x][x]}{0.50 - x}$$

If you rearrange the equation, it results in a quadratic equation (which we would like to avoid). However, since K_a is small, we can assume that x is negligible compared to 0.50, consequently, to simplify our calculation you can disregard the x in the E box under HA. If you know how to use the quadratic formula, you can do the calculation with the quadratic. (The answer will be the same as disregarding the x.)

Continue to use algebra to solve for x.

$$7.3\times10^{-8}\approx\frac{x^2}{0.50}$$

$$\left(7.3\times10^{-8}\right)(0.50)=\left(\frac{x^2}{0.50}\right)(0.50)$$

$$\sqrt{3.65\times10^{-8}}=\sqrt{x^2}$$

$$1.9\times10^{-4}=x$$

Step 3: Calculate equilibrium concentrations.

The concentration of the hydronium and A^- ions will be equal to x.

$[H_3O^+] = [A^-] = x = 1.9\times10^{-4}$ mol/L

$[HA] = 0.50 - x = 0.50 - 1.9\times10^{-4} = 0.49981$ mol/L

$[HA] = 0.50$ mol/L

The concentration of the HA is equal to 0.50 – x; however, when we subtract x from 0.50 mol/L, we get a value which is about equal to 0.50 mol/L.

This means our assumption that x is negligible is valid!

Percent Dissociation

You learned in the previous lesson that acids and bases can be described in terms of strong or weak, concentration, and degree of dissociation. The acid and base dissociation constants represent the degree of dissociation or ionization achieved by the acid or base. Another way to describe the amount of dissociation is by **percent dissociation**.

Percent dissociation is calculated in a manner similar to calculating your percentage on a test or an assignment. For percent dissociation, use the hydronium ion or hydroxide ion concentrations. (This will be explained later in this module.)

$$\text{percent dissociation} = \frac{\text{concentration of dissociated species}}{\text{original concentration of acid or base}} \times 100$$

For the acid HA,

$$\text{percent dissociation} = \frac{\left[H_3O^+\right]}{[HA]} \text{ x } 100$$

or, for the base B,

$$\text{percent dissociation} = \frac{\left[OH^-\right]}{[B]} \text{ x } 100$$

In the above expressions, [HA] and [B] are the initial concentrations of the acid and base, respectively. Obviously, your answers will be expressed as a percentage (%).

Percent Dissociation Problems

Example 4

Calculate the percent dissociation of a 0.100 mol/L solution of formic acid (HCOOH) if the hydronium ion concentration is 4.21×10^{-3} mol/L.

Step 1: Substitute concentrations into the percent dissociation equation.

$$\text{percent dissociation} = \frac{\left[H_3O^+\right]}{[HCOOH]} \times 100$$

$$= \frac{4.21 \times 10^{-3}}{0.100} \times 100 = 4.21\% \text{ dissociated}$$

Example 5

Calculate the K_b of the hydrogen phosphate ion (HPO_4^{2-}) if a 0.25 mol/L solution of sodium hydrogen phosphate is dissociated to 8%.

Step 1: Write out the dissociation equation.

$$HPO_4^{2-} + H_2O \rightleftharpoons H_2PO_4^- + OH^-$$

Step 2: Rearrange the percent dissociation equation to calculate hydroxide ion concentration. In order to calculate K_b, you need to find the concentration of the hydroxide and the dihydrogen phosphate ions.

$$\text{percent dissociation} = \frac{\left[\text{OH}^-\right]}{\left[\text{HPO}_4^{\ 2-}\right]} \times 100$$

$$\left[\text{OH}^-\right] = \frac{(\text{percent dissociation})\left(\text{HPO}_4^{\ 2-}\right)}{100}$$

$$= \frac{(8\%)(0.25 \text{ mol/L})}{100}$$

$$= 2.0 \times 10^{-2} \text{ mol/L}$$

$$\left[\text{OH}^-\right] = \left[\text{H}_2\text{PO}_4^{\ -}\right] = 2.0 \times 10^{-2} \text{ mol/L}$$

Step 3: Write the equilibrium law and substitute the values.

$$K_b = \frac{\left[\text{H}_2\text{PO}_4^{\ -}\right]\left[\text{OH}^-\right]}{\left[\text{HPO}_4^{\ 2-}\right]}$$

$$= \frac{\left(2.0 \times 10^{-2}\right)\left(2.0 \times 10^{-2}\right)}{0.25}$$

$$K_b = 1.6 \times 10^{-3}$$

pH, K_a, and K_b

Example 6

Calculate the pH of a 0.10 mol/L solution of the imaginary acid, HA, if its K_a is 4.0×10^{-9}.

In order to find the pH, you need to determine the hydronium ion concentration. To find the hydronium ion concentration when given K_a and the acid concentration, set up an ICE table.

Step 1: Set up an ICE table.

	HA	$+ H_2O$	$\rightleftharpoons$	H_3O^+	A^-
I	0.10			0.0	0.0
C	−x			+ x	+ x
E	0.10−x			x	x

Step 2: Substitute equilibrium values into the equilibrium law and solve for *x*.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$4.0 \text{ x } 10^{-9} = \frac{(x)(x)}{0.10 - x}$$

Since K_a is small, you can assume x is insignificant compared to 0.10 mol/L.

$$4.0 \times 10^{-9} = \frac{x^2}{0.10}$$

$$\left(4.0 \times 10^{-9}\right)(0.10) = \left(\frac{x^2}{\cancel{0.10}}\right)(\cancel{0.10})$$

$$\sqrt{4.0 \times 10^{-10}} = \sqrt{x^2}$$

$$2.0 \times 10^{-5} = x$$

$$[H_3O^+] = x = 2.0 \times 10^{-5} \text{ mol/L}$$

Step 3: Calculate the pH.

$$pH = -\log[H_3O^+] = -\log(2.0 \times 10^{-5})$$

$$pH = 4.70$$

Example 7

The pH of a 0.20 mol/L solution of CO_3^{2-} ions is 10.75. What is the K_b of carbonate ions?

Step 1: Write the dissociation equation.

$$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$$

The procedure to solve this problem is the opposite of the previous example. In order to find the K_b, you must determine the hydroxide ion concentration, since $[HCO_3^-] = [OH^-]$. Remember, you can get the hydroxide ion concentration from the pH.

Step 2: Use pH to calculate $[OH^-]$. First find the hydronium ion concentration.

$$[H_3O^+] = 10^{-pH} = 10^{-10.75} = 1.78 \times 10^{-11} \text{ mol/L}$$

Then use K_w to find hydroxide ion concentration.

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \text{ x } 10^{-14}}{1.78 \text{ x } 10^{-11}} = 5.6 \text{ x } 10^{-4} \text{ mol/L}$$

Step 3: Substitute into equilibrium law.

$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]} = \frac{(5.6 \text{ x } 10^{-4})(5.6 \text{ x } 10^{-4})}{0.20} = 1.6 \text{ x } 10^{-6}$$

Learning Activity 5.6: Problem Solving with K_a and K_b

Solve the following problems. Be sure to show all your work and use correct units for your answers. If needed, use values from the K_a table that can be found in the appendices at the end of the course.

1. Calculate the pH of a 0.10 mol/L solution of hypochlorous acid, HOCl ($K_a = 3.5 \text{ x } 10^{-8}$)
2. Calculate the K_a of a weak acid, HX, if a 0.25 mol/L solution has a pH of 4.40.
3. Calculate the concentration of a solution of acetic acid ($K_a = 1.8 \text{ x } 10^{-5}$), which has a pH = 3.00.
4. Hydrazine (N_2H_4) is a weak base with a K_b of $3.0 \text{ x } 10^{-6}$. The reaction of hydrazine in water is $H_2NNH_2 + H_2O \rightleftharpoons H_2NNH_3^+ + OH^-$. Calculate the pH of a 2.0 mol/L solution of hydrazine.
5. Calculate the percent dissociation of a 0.20 mol/L solution of the weak acid, HNO_2, if the pH of the solution is 4.20.

Check the answer key.

Lesson Summary

In this lesson, you practiced problem-solving methods that involved pH, K_a, K_b, and percent dissociation. Some of these problems required you to organize an ICE table like you did in the equilibrium module. In the next lesson, you will calculate the concentration of an unknown base or acid.

Assignment 5.6: pH and Percent Dissociation (15 marks)

Solve the following problems. Be sure to show all your work and use correct units for your answers. If needed, use values from the K_a table that can be found in Appendix D at the end of the course.

1. A 0.20 mol/L solution of the weak base HPO_4^{2-} has a pH of 9.00. Find the K_b. *(7 marks)*

continued

Assignment 5.6: pH *and* Percent Dissociation (continued)

2. At 25°C, a 0.010 mol/L ammonia solution is 4.3% ionized. Calculate the pH. *(8 marks)*

Lesson 7: Acid-Base Concentration (2 hours)

Lesson Focus

SLO C12-5-10: Perform a laboratory activity to determine the concentration of an unknown acid or base, using a standardized acid or base.

Lesson Introduction

In the previous lesson, you were introduced to problem-solving methods that involved pH, K_a, K_b, and percent dissociation. In this lesson, you will learn about titration and how to determine the concentration of an unknown base or acid.

Titration

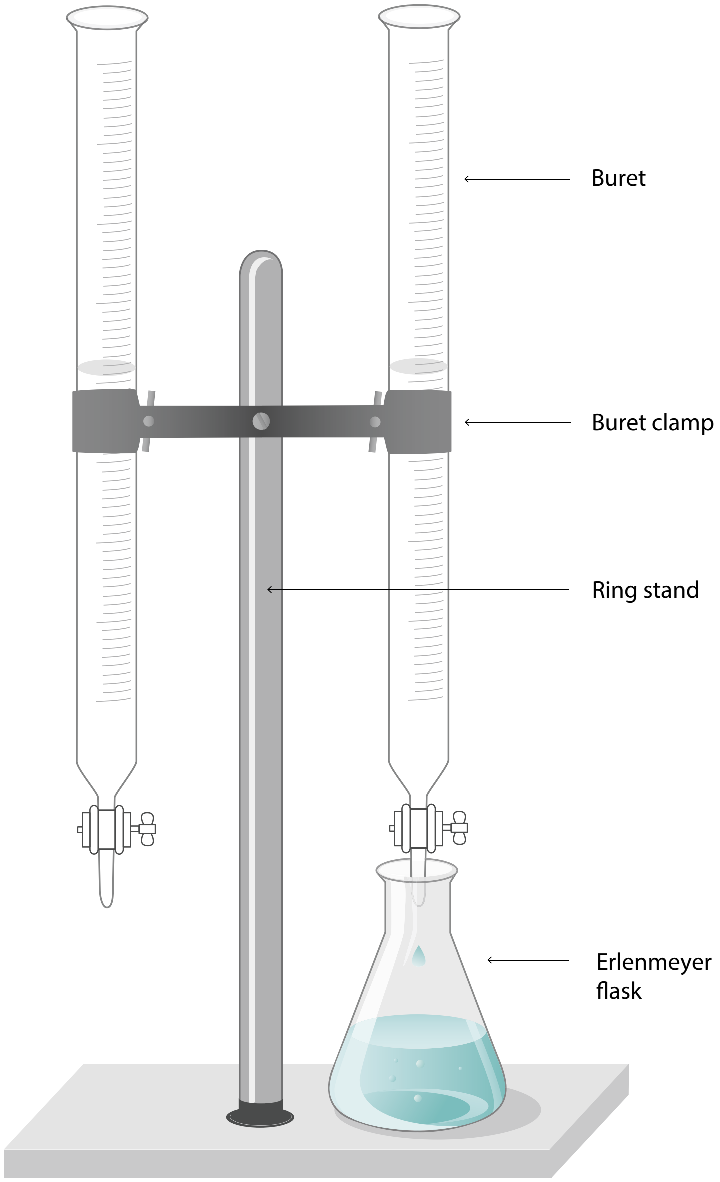
A **titration** is a carefully controlled neutralization process used to determine an unknown concentration. To perform a titration, a standard solution is required. A **standard solution** is a solution of a strong acid or base of **known concentration**. To determine the concentration of an unknown acid, a basic standard solution is required, and an acid standard solution is required for an unknown base. The standard solution is usually added to a sample of unknown concentration until neutralization has occurred.

The picture on the following page shows the titration apparatus that you would assemble in the lab to carry out a titration. A **burette** is used to control the additions of acids and bases in a titration. The buret readings are used to determine the volumes of acids and bases added in a titration.

The following video shows how and why titrations are performed:

www.youtube.com/watch?v=i7jnSaf1Muc

If you want to see other examples, use a search engine such as www.google.ca/.



In the first module of this course, you learned that when an acid or base is at the exact point where it is neutralized, the moles of hydronium ions from the acid and the moles of hydroxide ions from the base are equivalent. The point at which the amount of standard acid or base solution added neutralizes the unknown sample is called the **equivalence point**. The equivalence point can be determined by adding a pH indicator to the unknown sample or by using a pH meter to measure the pH change as standard solution is added to the sample. If an indicator is used, the point in the titration at which the desired colour forms is called the **end point** of the titration. The indicator is chosen so that the end point and the equivalence point are very close. The end point and the equivalence point are **not** the same thing.

Performing a Titration

When performing an acid-base titration on an unknown solution, both the standard and unknown solutions are placed into separate burets. For this example, we will determine the concentration of a hydrochloric acid solution (unknown) using a 0.100 mol/L solution of NaOH.

Step 1: Fill each burette and take the initial volume readings from the bottom of the meniscus (the smile-shaped surface of the solution in the container).

Step 2: Release a sample of the unknown acid solution into an Erlenmeyer flask. Read the acid volume in the burette (this is the final acid volume).

Note

The volume of the unknown acid sample is equal to the following:

final volume reading—initial volume reading

Step 3: Add an indicator, like phenolphthalein. The end point will be reached when the sample turns a light pink. Phenolphthalein turns pink at a pH of about 8.2.

Step 4: Slowly add the standard base solution to the flask, while stirring or swirling, until the solution in the flask turns a light pink and stays pink, in the case of phenolphthalein. Take the reading on the base burette and call this the final base volume.

Step 5: Repeat this procedure until there are several trials reflecting consistent data. This data will be used to determine the unknown concentration of the acid.

Using Titrations to Determine an Unknown Concentration

Once you have determined the volume of unknown acid required to neutralize a given volume of a base of known concentration, you can then determine the concentration of the acid. In the previous section, you read the detailed steps for a titration that was performed using a standard solution of 0.100 mol/L NaOH and an unknown HCl solution. Here is a sample of data that might be collected from a neutralization.

	Base	**Acid**
Final Burette Reading	14.45 mL	12.57 mL
Initial Burette Reading	0.62 mL	1.13 mL

Using this information, you can calculate the concentration of the acid.

Step 1: Determine the volume of acid and base used by subtracting the initial from the final readings:

	Base	**Acid**
Final Burette Reading	14.45 mL	12.57 mL
Initial Burette Reading	0.62 mL	1.13 mL
Volume (mL)	13.83 mL	11.44 mL
Volume (L)	0.01383 L	0.01144 L

Step 2: Find the moles of unknown solution. The neutralization reaction is

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

At the end point, moles H+ = moles OH– because the ratio is 1:1 between HCl and NaOH. Next, calculate the moles of base used.

$$\text{moles NaOH} = C \times V = (0.100\ \text{mol/L})(0.01383\ \text{L}) = 0.001383\ \text{moles}$$

The sample of acid contains 0.001383 moles.

Step 3: Find the concentration.

$$C = \frac{\text{mol}}{\text{L}} = \frac{0.001383\ \text{mol}}{0.01144\ \text{L}} = 0.121\ \text{mol/L}$$

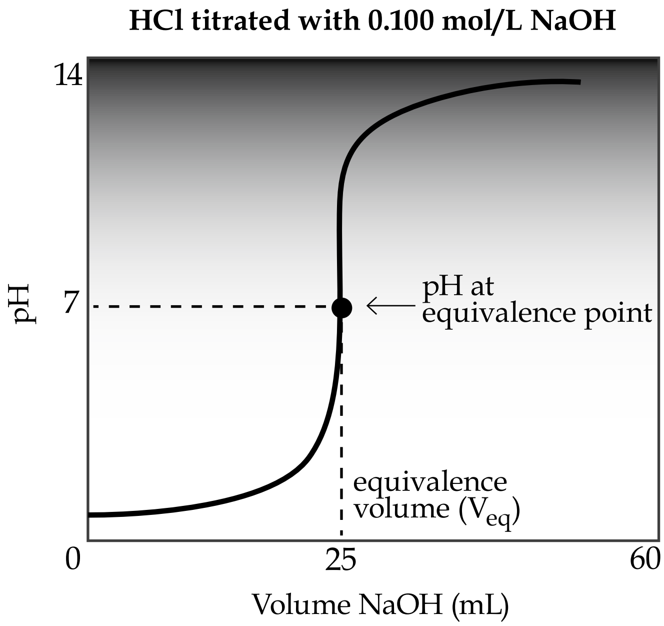
The concentration of the unknown acid is 0.121 mol/L.

Titration Curves

A titration reaction can be graphed. The data plotted would be volume (of acid or base) versus pH changes. In this example, 25.0 mL of an HCl solution is titrated with a 0.100 mol/L solution of NaOH. The equation for this reaction is as follows:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

At the equivalence point, the hydronium ion and hydroxide ion concentrations are equivalent, as are the sodium and chloride ion concentrations. The solution is neutral (at 25°C the pH will be 7.0). The expected titration curve would appear as shown on the graph on the following page if you graph the change in pH as base is added.



Note

The equivalence point is found in the centre of the steepest portion of the curve (where the curve appears to change directions).

The equivalence volume is 25.0 mL of 0.100 mol/L NaOH, since the acid:base stoichiometry is 1:1,

$$\text{moles } H^+ = \text{moles } OH^- = (0.100 \text{ mol/L})(0.025 \text{ L}) = 0.0025 \text{ moles}$$

$$[HCl] = 0.0025 \text{ moles} \div 0.025 \text{ L} = 0.100 \text{ mol/L HCl}$$

The grey shading on the graph indicates the colour change of phenolphthalein as the pH changes. Despite the colour change at a pH higher than equivalence, the difference is not significant because the curve is so steep at the equivalence point.

The following web link provides definitions and additional explanations of titration and the related terms involved. This may help you better understand the concept of titration. See www.sparknotes.com/chemistry/acidsbases/titrations/section1.html.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Learning Activity 5.7: Calculating an Unknown Concentration

1. Write the complete set of reactions that occur when the following acid and bases are reacted. Look back at Lesson 2 if you need some review first.

   a. Hydrochloric acid and sodium hydroxide

      Molecular: ____

      Ionic: ____

      Net Ionic: ____

   b. Sulfuric acid and potassium hydroxide

      Molecular: ____

      Ionic: ____

      Net Ionic: ____

   c. Nitric acid and calcium hydroxide

      Molecular: ____

      Ionic: ____

      Net Ionic: ____

   d. Phosphoric acid and lithium hydroxide

      Molecular: ____

      Ionic: ____

      Net Ionic: ____

   e. Sulfuric acid and aluminum hydroxide

      Molecular: ____

      Ionic: ____

      Net Ionic: ____

continued

Learning Activity 5.7: Calculating an Unknown Concentration (continued)

2. Calculate the concentration of a solution of KOH if 140.0 mL of the base is needed to neutralize 22.5 mL of 0.175 mol/L acid HNO_3.
3. Determine the volume of 0.750 mol/L LiOH required to neutralize 56.0 mL of 0.350 mol/L H_2SO_4.
4. Calculate the concentration of H_3PO_4 if 25.0 mL is required to neutralize 19.0 mL of 0.830 mol/L KOH.
5. What volume of 0.100 mol/L HCl is required to just neutralize 10.0 g of solid NaOH?
6. What volume of 0.250 mol/L nitric acid is needed to just neutralize 25.0 g of magnesium hydroxide?
7. What mass of calcium hydroxide is needed to just neutralize 250. mL of 0.100 mol/L HCl? Assume the volume does not change.

Check the answer key.

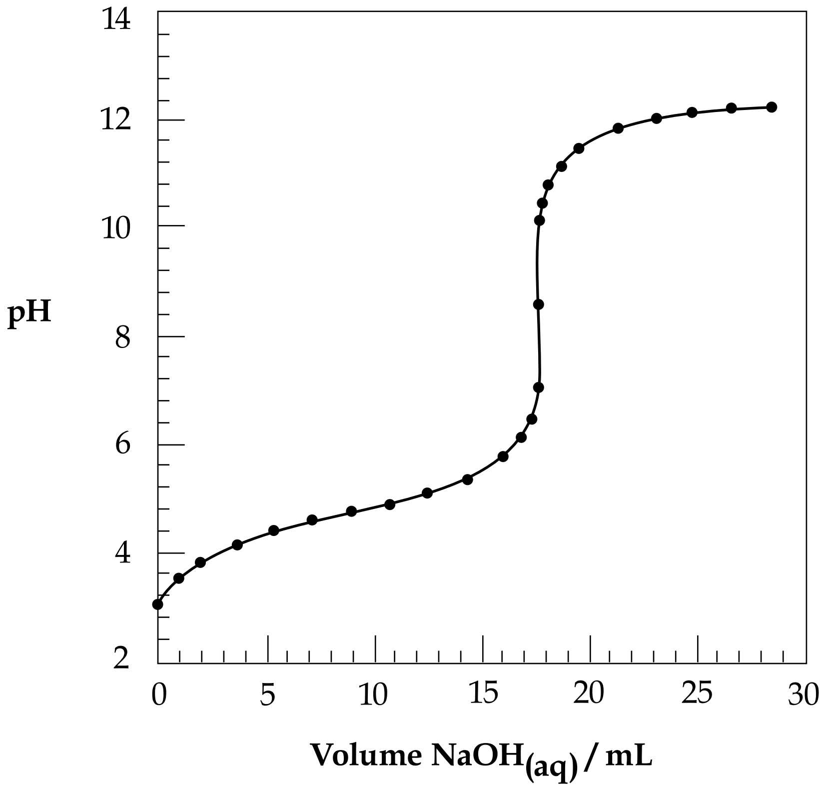
Lesson Summary

In this lesson, you learned about the process called acid-base titration. You also practiced calculating the concentration of an unknown acid or base using titration data. In the next lesson, you will predict whether an aqueous solution of a given ionic compound will be acidic, basic, or neutral, given the formula.

Notes

Assignment 5.7: Titration Curves (16 marks)

Answer the following questions, showing **all** your work. **Use the titration curve provided to help you answer questions 1 through 4.**



A student weighed 0.225 g of a solid unknown acid into a flask, added about 50 mL of water, and then titrated the resulting mixture with a standard 0.1245 mol/L solution of aqueous sodium hydroxide, measuring the pH of the mixture after each addition. The above graph shows the titration curve obtained. Using this graph, answer the following questions:

1. What would be a suitable indicator for this solution? Why? The pH range for colour change chart from Lesson 4 may assist you with this question. *(2 marks)*

___

___

___

continued

Assignment 5.7: Titration Curves (continued)

2. Estimate the molar mass of the unknown acid, assuming the unknown acid is a weak monoprotic acid. *(4 marks)*

3. Would the results obtained have been different had the student added 100 mL of water? Explain your answer. *(1 mark)*

4. What happens to the graph as the titration nears its equivalence point? *(1 mark)*

5. What volume of 0.240 mol/L H_2SO_4 can be neutralized by 50.0 mL of 0.360 mol/L NaOH? *(8 marks)*

LESSON 8: ACID-BASE PREDICTIONS (2 HOURS)

Lesson Focus

SLO C12-5-11: Predict whether an aqueous solution of a given ionic compound will be acidic, basic, or neutral, given the formula.

Lesson Introduction

In this lesson, you will learn to appreciate that salts can be something other than neutral. You might recall that an acid combines with a base in a neutralization reaction, and a salt and water are produced. However, the resulting aqueous salt solution can be neutral, acidic, or basic, depending on the strength of the acid and base that are reacted. In this lesson, you will predict whether an aqueous solution of a given ionic compound will be acidic, basic, or neutral, given the formula.

Hydrolysis

You may recall having learned that a **neutralization** reaction occurs between an acid and a base. The result of such a reaction is the production of a salt and water. During neutralization reactions, a solution with a neutral pH may not be the result. Even though the acid and base neutralize each other, the salt produced may affect the pH of the solution.

A salt can react with the water to produce acidic or basic solutions. This process is called salt **hydrolysis**. When a salt dissociates, one of two scenarios may occur. First, the anions of the dissociated salt may accept hydrogen ions from the water, thus producing a **basic solution**. Alternatively, the cations of the dissociated salt may donate hydrogen ions to the water, thus producing an **acidic solution**.

There are three possible neutralization reactions and resulting salts.

Example A: Strong acid + Strong base

The cation from a strong base plus the anion from a strong acid result in a neutral solution (pH = 7).

Type of Salt	Examples of Salt Produced	Ions That Undergo Hydrolysis	pH
Cation from strong base Anion from strong acid	NaCl, KI, KNO_3, RbBr $BaCl_2$	None	≈7

Example B: Weak acid + Strong base

The cation from a strong base plus the anion from a weak acid produce a solution (pH > 7).

Type of Salt	Examples of Salt Produced	Ions That Undergo Hydrolysis	pH
Cation from strong base Anion from weak acid	$NaC_2H_3O_2$, KNO_2	Anion	>7

Example C: Strong Acid + Weak Base

The cation from a weak base plus the anion from a strong acid produce a solution (pH < 7).

Type of Salt	Examples of Salt Produced	Ions That Undergo Hydrolysis	pH
Cation from weak base Anion from strong acid	NH_4Cl, NH_4NO_3	Cation	<7

Neutral Salts

When strong acids and strong bases react, a **neutral salt** is produced, as explained in Example A on the previous page. For example,

$$HCl + NaOH \rightarrow NaCl + H_2O$$

NaCl contains a group 1 ion (sodium) and the chloride ion. The chloride ion is the conjugate base of a strong acid. Since it is the conjugate of a strong acid, it has a very low attraction for protons. In fact, water has a higher proton attraction than chloride ions; therefore, in water, chloride ions do not affect the pH.

The following anions of the strong acids will not affect the pH of a solution:

- chloride
- perchlorate
- iodide
- bromide
- nitrate

> **Note**
>
> Any group 1 or 2 spectator ions will not affect the pH of a solution.

Basic Salts

Basic salts result from the reaction of a strong base with a weak acid, as outlined in Example B on the previous page. The basic solution is due to the presence of the conjugate base of the weak acid. For example, sodium acetate, $NaC_2H_3O_2$, is formed from the reaction of sodium hydroxide and acetic acid, a weak acid. In water, sodium acetate dissociates according to the following reaction:

$$NaC_2H_3O_{2(aq)} \rightarrow Na^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

Sodium acetate dissolves in water to produce sodium cation ions and acetate anions. The sodium ions do not react with water because they are group 1 ions, but the acetate ion will react with water because it is the conjugate base of a weak acid. In water,

$$C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HC_2H_3O_{2(aq)} + OH^-{}_{(aq)}$$

Since hydroxide ions are produced, the solution is basic.

Acidic Salts

Acidic salts are due to the formation of the conjugate acid of a weak base, shown as Example C. For example, NH_4NO_3 is the salt formed from the reaction of ammonia and nitric acid. In water, NH_4NO_3 dissociates according to the following chemical equation:

$$NH_4NO_{3(aq)} \rightarrow NH_4{}^{+}{}_{(aq)} + NO_3{}^{-}{}_{(aq)}$$

The nitrate ion will not react with water because it is the negative ion of a strong acid. You may recognize the ammonium ion, $NH_4{}^{+}$, as the conjugate acid of NH_3. When the ammonium ion is in water, it reacts with the water as an acid.

$$NH_4{}^{+}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^{+}{}_{(aq)}$$

Since H_3O^+ is produced, the salt is, therefore, acidic.

Example 1

Is potassium carbonate (K_2CO_3) acidic, basic, or neutral?

Step 1: Identify the ions present after dissociation occurs.

Potassium carbonate will dissociate in water as follows:

$$K_2CO_{3(s)} \rightarrow 2\ K^{+}{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$$

The ions present are K^+ and $CO_3{}^{2-}$.

Step 2: Determine if each ion will affect the pH.

The potassium ion is a group 1 ion so it will not affect the pH; it is neutral in water.

The carbonate ion is the conjugate base of the weak acid $HCO_3{}^-$ (or the weak acid H_2CO_3). In water,

$$CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^{-}{}_{(aq)} + OH^{-}{}_{(aq)}$$

The carbonate ions react to form an excess of hydroxide ions; as such, the solution is basic.

You can also use the K_a values of the solution to help you predict the nature of the resulting salt.

Example 2

Is sodium acetate ($NaC_2H_3O_2$) acidic, basic, or neutral?

Step 1: Identify the ions present after dissociation occurs.

Sodium acetate solid dissolves in water to produce sodium cation ions and acetate anions.

$$NaC_2H_3O_2 \xrightarrow{H_2O} Na^+_{(aq)} + C_2H_3O^-_{2(aq)}$$

$Na^+_{(aq)} + H_2O_{(l)} \rightarrow$ no reaction because Na^+ is a spectator ion

Step 2: Use the K_a value to help you predict the outcome.

Because the K_a for $HC_2H_3O_2$ is very small (1.8×10^{-5}), the reaction below tends to go forward as written to remove hydrogen ions from solution, and it leaves an excess of hydroxide ions.

$$C_2H_3O^-_{2(aq)} + H_2O_{(l)} \rightarrow HC_2H_3O_{2(aq)} + OH^-_{(aq)}$$

The acetate ions react to form an excess of hydroxide ions; therefore, the solution is basic.

Example 3

Determine whether sodium sulphate salt (Na_2SO_4) is acidic, basic, or neutral in water. Provide a rationale to support your prediction, including the ions formed in each reaction.

Ions present: N^+, SO_4^{2-}
Na^+: group 1 ion, therefore neutral
SO_4^{2-}: conjugate base of the weak acid HSO_4^-, therefore basic
$SO_4^{2-} + H_2O \leftrightarrow HSO_4^- + OH^-$
Na_2SO_4 is basic in water.

Example 4

Determine whether potassium phosphate salt (K_3PO_4) is acidic, basic, or neutral in water. Provide a rationale to support your prediction, including the ions formed in each reaction.

Ions present: K^+, PO_4^{3-}
K^+: group 1 ion; therefore, neutral
PO_4^{3-}: conjugate base of the weak acid HPO_4^{2-}, therefore basic
$PO_4^{3-} + H_2O \leftrightarrow HPO_4^{2-} + OH^-$
K_3PO_4 is basic in water.

Learning Activity 5.8: Predicting the Nature of Salts

Determine whether each of the following salts is acidic, basic, or neutral in water. For each answer, give a rationale to support your prediction, including the ions formed in each reaction.

1. Sodium nitrate
2. Sodium sulphide
3. Potassium chloride
4. Magnesium nitrate
5. Sodium carbonate
6. Ammonium iodide
7. Potassium acetate
8. Calcium oxide

Check the answer key.

Lesson Summary

In this lesson, you learned how to predict whether a salt formed during a neutralization reaction is acidic, basic, or neutral. This is the last lesson in Module 5.

Assignment 5.8: Acid-Base Predictions (5 marks)

1. Determine whether sodium nitrite salt ($NaNO_2$) is acidic, basic, or neutral in water. Provide a rationale to support your prediction, including the ions present in each reaction.

Notes

MODULE 5 SUMMARY

In this module, you learned that acids and bases are two important classes of chemical compounds. You encounter acids and bases everyday in foods and items such as personal care products (for example, soap and toothpaste), medications, and many other products. In addition, acids and bases are very important to the manufacturing industry in the production of dyes and textiles, paper, polymers, and fertilizers, just to name a few.

There are acids and bases at work in your body, as well. For example, body systems are very sensitive to the acidity of your blood. Your stomach secretes hydrochloric acid to aid in digestion. You might occasionally take an antacid to control the amount of acid in your stomach.

Acids and bases play a large part in our everyday lives. In this module, you defined acids and bases and studied acid-base reactions.

You have now completed Module 5. Read the following directions and double-check that you have submitted all the required work.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 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 5 assignments and organize your material in the following order:

☐ Cover Sheet for Module 5 (found at the end of the course Introduction)
☐ Assignment 5.1: Acid-Base Theories
☐ Assignment 5.2: Identifying Conjugate Pairs
☐ Assignment 5.3: Solving K_w Problems
☐ Assignment 5.4: Solving pH Problems
☐ Assignment 5.5: Acid and Base Equilibrium
☐ Assignment 5.6: pH and Percent Dissociation
☐ Assignment 5.7: Titration Curves
☐ Assignment 5.8: Acid-Base Predictions

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

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GRADE 12 CHEMISTRY (40S)

Module 5
Acids and Bases

Learning Activity Answer Keys

MODULE 5:
ACIDS AND BASES

Learning Activity 5.1: Acid-Base Definitions

1. Compare the three definitions of acids and bases that were introduced in this lesson.

Theory	Acid	Base
Arrhenius	**H^+ producer**	**OH^- producer**
Brønsted-Lowry	**H^+ donor**	**H^+ acceptor**
Lewis	**Electron-pair acceptor**	**Electron-pair donor**

3. Match the terms with the correct description below.

d	1. It is a substance that can accept a pair of electrons.	a. Amphoteric
a	2. It describes a substance that can act as both an acid and a base.	b. Hydroxide ion
b	3. OH^-	c. Hydronium ion
c	4. H_3O^+	d. Lewis acid
e	5. It is the formula of the hydrogen ion.	e. H^+

5. What is the Brønsted-Lowry definition of an acid? Of a base?

   A Brønsted-Lowry acid donates H^+ ions while a Brønsted-Lowry base accepts H^+ ions.

6. Identify the Lewis acid and the Lewis base in the following reaction:

   $H^+ + Cl^- \rightleftharpoons HCl$

   Lewis Acid: H^+ (accepts a pair of electrons)

   Lewis base: Cl^- (donates a pair of electrons)

Learning Activity 5.2: Writing Balanced Acid-Base Reactions

1. Complete the following acid-base reactions, using the examples provided in the lesson to help you predict the products and assure conservation of charges. Then, identify which of these are Arrhenius acids and bases and which are Brønsted-Lowry acids and bases.

   a. $HI_{(aq)} + H_2O_{(l)} \rightarrow$

      $\mathbf{HI_{(aq)} + H_2O_{(l)} \rightarrow I^-_{(aq)} + H_3O^+_{(aq)}}$
      Brønsted-Lowry and Arrhenius acid

   b. $HF_{(aq)} + H_2O_{(l)} \rightarrow$

      $\mathbf{HF_{(aq)} + H_2O_{(l)} \rightarrow F^-_{(aq)} + H_3O^+_{(aq)}}$
      Brønsted-Lowry and Arrhenius acid

   c. $C_2H_3O_2{}^-{}_{(aq)} + H_2O_{(l)} \rightarrow$

      $\mathbf{C_2H_3O_2{}^-{}_{(aq)} + H_2O_{(l)} \rightarrow HC_2H_3O_{2(aq)} + OH^-_{(aq)}}$
      Brønsted-Lowry base only

   d. $CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow$

      $\mathbf{CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow HCO_3{}^-{}_{(aq)} + OH^-_{(aq)}}$
      Brønsted-Lowry base only

   e. $HSO_3{}^-{}_{(aq)} + H_2O_{(l)} \leftrightarrow$

      $\mathbf{HSO_3{}^-{}_{(aq)} + H_2O_{(l)} \rightarrow SO_3{}^{2-}{}_{(aq)} + H_3O^+_{(aq)}}$
      and/or

      $\mathbf{HSO_3{}^-{}_{(aq)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)} + OH^-_{(aq)}}$
      Brønsted-Lowry and Arrhenius acid

2. Since Brønsted-Lowry acids are proton donors, what do you think the term *triprotic acid* means?

   A triprotic acid is one that has three hydrogen ions to donate, such as $\mathbf{H_3PO_4}$.

3. What does the Brønsted-Lowry acid-base theory state in regards to water?

   Water acts as a base when it accepts a hydrogen ion and acts as an acid when it donates a hydrogen ion; therefore, water is amphoteric.

Learning Activity 5.3: K_W Calculations

1. If the [OH⁻] in a sodium hydroxide solution is 0.050 mol/L, what is $[H_3O^+]$?

$$K_W = [H_3O^+][OH^-];\ \left[\mathbf{H_3O^+}\right] = \frac{K_W}{\left[\mathbf{OH^-}\right]} = \frac{\mathbf{1.0 \times 10^{-14}}}{\mathbf{0.050\ mol/L}} = \mathbf{2.0 \times 10^{-13}\ mol/L\ H_3O^+}$$

2. 0.250 mole of hydrogen chloride gas is dissolved in 2.00 L of water. Write the dissociation equation for this gas and calculate both $[H_3O^+]$ and [OH⁻].

$$HCl_{(g)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$\mathbf{[HCl] = \frac{mol}{V} = \frac{0.250\ moles}{2.00\ L} = 0.125\ mol/L}$$

HCl is a strong acid so $\left[\mathbf{H_3O^+}\right] = \mathbf{0.125\ mol/L}$

$$K_W = [H_3O^+][OH^-];\ \left[\mathbf{OH^-}\right] = \frac{\mathbf{K_W}}{\left[\mathbf{H_3O^+}\right]} = \frac{\mathbf{1.0 \times 10^{-14}}}{\mathbf{0.125\ mol/L}} = \mathbf{8.0 \times 10^{-14}\ mol/L\ OH^-}$$

3. 10.0 g of lithium hydroxide is dissolved in 750.0 mL of water. Write the dissociation equation and calculate both $[H_3O^+]$ and [OH⁻].

$$\mathbf{LiOH \rightarrow Li^+ + OH^-}$$

$$\mathbf{LiOH = 23.9\ g/mol}$$

$$\mathbf{moles\ LiOH = \frac{mass}{M.M.} = \frac{10.0\ g}{23.9\ g/mol} = 0.4184\ moles\ LiOH}$$

$$\mathbf{C = \frac{mol}{V} = \frac{0.418\ mol}{0.7500\ L} = 0.557\ mol/L}$$

$$\mathbf{\left[OH^-\right] = [LiOH] = 0.557\ mol/L}$$

$$K_W = [H_3O^+][OH^-];\ \left[\mathbf{H_3O^+}\right] = \frac{K_W}{\left[\mathbf{OH^-}\right]} = \frac{\mathbf{1.0\ x\ 10^{-14}}}{\mathbf{0.557\ mol/L}} = \mathbf{1.80 \times 10^{-14}\ mol/L\ H_3O^+}$$

4. 10.0 g of calcium hydroxide is dissolved in 400.0 mL of solution. Write the dissociation equation and calculate both [H_3O^+] and [OH^-].

$$\mathbf{Ca(OH)_{2(s)} \rightarrow Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}}$$

$$\mathbf{Ca(OH)_2 = 74.1\ g/mol}$$

$$\mathbf{moles\ Ca(OH)_2 = \frac{m}{C} = \frac{10.0\ g}{74.1\ g/mol} = 0.135\ moles\ Ca(OH)_2}$$

$$\mathbf{C_{Ca(OH)2} = \frac{mol}{V} = \frac{0.135\ mol}{0.400\ L} = 0.3375\ mol/L}$$

$$\mathbf{\left[OH^-\right] = 2 \times \left[Ca(OH)_2\right] = 2 \times 0.3375\ mol/L = 0.675\ mol/L\ OH^-}$$

$$K_W = [H_3O^+][OH^-]; \mathbf{\left[H_3O^+\right] = \frac{K_W}{\left[OH^-\right]} = \frac{1.0\ x\ 10^{-14}}{0.675\ mol/L} = 1.48 \times 10^{-14}\ mol/L\ H_3O^+}$$

5. If the [H_3O^+] of a barium hydroxide solution is 1.00×10^{-13} mol/L, calculate the [OH^-]. How many grams of barium hydroxide are needed to make a litre of this solution?

$$\mathbf{Ba(OH)_{2(s)} \rightarrow Ba^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}}$$

$$K_W = [H_3O^+][OH^-]; \mathbf{\left[OH^-\right] = \frac{K_W}{\left[H_3O^+\right]} = \frac{1.0 \times 10^{-14}}{1.00 \times 10^{-13}} = 0.10\ mol/L\ OH^-}$$

$$\mathbf{C_{Ba(OH)2} = \frac{\left[OH^-\right]}{2} = \frac{0.10\ mol/L}{2} = 0.050\ mol/L\ Ba(OH)_2}$$

$$\mathbf{molesBa(OH)_2 = C \times V = (0.050\ mol/L)(1.0\ L) = 0.050\ moles\ Ba(OH)_2}$$

$$\mathbf{mass = mol \times M = (0.050\ moles)(171.3\ g/mol) = 8.6\ g\ Ba(OH)_2}$$

6. Calculate the [H_3O^+] in milk of magnesia (magnesium hydroxide) that has an [OH^-] of 1.43×10^{-4} mol/L.

$$\mathbf{Mg(OH)_{2(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2\ OH^-{}_{(aq)}}$$

$$K_W = [H_3O^+][OH^-]; \mathbf{\left[H_3O^+\right] = \frac{K_W}{\left[OH^-\right]} = \frac{1.0 \times 10^{-14}}{1.43 \times 10^{-4}} = 6.99 \times 10^{-11}\ mol/L\ H_3O^+}$$

Learning Activity 5.4: The pH Scale

Part A: Calculations and Short Answer

1. Determine the pH of each of the following values
   a. $[H^+] = 2.0 \times 10^{-3}$

   $\mathbf{pH = -log[H^+] = -log(2.0 \times 10^{-3}) = 2.7}$

   b. $[H^+] = 3.5 \times 10^{-4}$

   $\mathbf{pH = -log[H^+] = -log(3.5 \times 10^{-4}) = 3.4}$

   c. $[OH^-] = 7.5 \times 10^{-7}$

   $\mathbf{pOH = -log[OH^-] = -log(7.50 \times 10^{-7}) = 6.1}$

   $\mathbf{pH = 14 - pOH = 14 - 6.1 = 7.9}$

   d. $[OH^-] = 9.0 \times 10^{-9}$

   $\mathbf{pOH = -log[OH^-] = -log(9.0 \times 10^{-9}) = 8.0}$

   $\mathbf{pH = 14 - pOH = 14 - 8.05 = 6.0}$

2. Determine the concentration of $[H^+]$ for each of the following.
   a. $pH = 7.0$

   $\mathbf{[H_3O^+] = 10^{-pH} = 10^{-7} = 1.0 \times 10^{-7}\ mol/L\ H_3O^+}$

   b. $pOH = 5.0$

   $\mathbf{pH = 14 - pOH = 14 - 5.0 = 9.0}$

   $\mathbf{[H_3O^+] = 10^{-pH} = 10^{-9} = 1.0 \times 10^{-9}\ mol/L\ H_3O^+}$

   c. $pH = 13.0$

   $\mathbf{[H_3O^+] = 10^{-pH} = 10^{-13} = 1.0 \times 10^{-13}\ mol/L\ H_3O^+}$

   d. $pOH = 11.0$

   $\mathbf{pH = 14 - pOH = 14 - 11.0 = 3.0}$

   $\mathbf{[H_3O^+] = 10^{-pH} = 10^{-3} = 1.0 \times 10^{-3}\ mol/L\ H_3O^+}$

3. Determine the pOH of the following if the pH is given.
   a. $pH = 5.4$

   $\mathbf{pOH = 14 - pH = 14 - 5.4 = 8.6}$

   b. $pH = 8.6$

   $\mathbf{pOH = 14 - pH = 14 - 8.6 = 5.4}$

4. Determine the concentration of the OH^- ions in the solutions with the following pH values.

   a. pH = 1.6

      pOH = 14 – pH = 14 – 1.6 = 12.4

      $[OH^-] = 10^{-pOH} = 10^{-12.40} = 4.0 \times 10^{-13}$ mol/L OH^-

   b. pH = 9.2

      pOH = 14 – pH = 14 – 9.2 = 4.8

      $[OH^-] = 10^{-pOH} = 10^{-4.80} = 1.6 \times 10^{-5}$ mol/L OH^-

5. Calculate the pH of each of the following solutions.

   a. 0.0020 mol/L HCl

      $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

      $[H_3O^+] = 0.0020$ mol/L

      $pH = -\log[H_3O^+] = -\log(0.0020) = 2.7$

   b. 4.00×10^{-4} mol/L $Ba(OH)_2$

      $Ba(OH)_2 \rightarrow Ba^{2+} + 2\,OH^-$

      $[OH^-] = 2(4.00 \times 10^{-4}$ mol/L$) = 8.0 \times 10^{-4}$ mol/L

      $$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{8.0 \times 10^{-4}} = 1.25 \times 10^{-11} \text{ mol/L } H_3O^+$$

      $pH = -\log[H_3O^+] = -\log(1.25 \times 10^{-11}) = 10.9$

6. Which indicator changes colour in a solution that has a pH of 2? Use the table provided in the lesson to help you answer this question.

   Thymol blue

7. Which indicator(s) would you choose to show that a solution has changed from a pH of 3 to a pH of 5? From a pH of 4.6 to a pH of 4.9?

   pH of 3 to a pH of 5: bromphenol blue or bromcresol blue

   pH of 4.6 to a pH of 4.9: bromcresol blue or methyl red

8. Red cabbage juice can be used as an indicator to show the relative pH ranges of different solutions. Use the data that you collected previously as well as your knowledge of acids and bases to answer each of the following questions:

   a. CLR is a commonly used household cleaner. A student tests CLR in a solution of red cabbage juice and the solution turns a bright red colour. What would be the relative pH range of CLR?

      0–3

b. A clear, carbonated beverage is tested in red cabbage juice and the solution turns a light purple colour. What would be the pH range of the beverage?

4–6

c. An antacid tablet is added to red cabbage juice and the solution turns a light green colour. What would be the pH range of the antacid?

8–11

9. Are there any OH^- ions in an acidic solution with a pH of 3.25? How do you know?

Yes, there are always some hydroxide ions in an aqueous solution. This is evident since the product of the hydrogen ion concentration and the hydroxide ion concentration must always equal the K_W.

Part B: Multiple Choice

1. Which of these solutions is the **most** basic?

a. $[H^+] = 1 \times 10^{-2}$

b. $[OH^-] = 1 \times 10^{-4}$

c. $[H^+] = 1 \times 10^{-11}$

d. $[OH^-] = 1 \times 10^{-13}$

2. In a neutral solution, the $[H^+]$ is

a. 10^{-14}

b. 0

c. 1×10^{7} mol/L

d. Equal to $[OH^-]$

3. A solution in which the hydroxide ion concentration is 1×10^{-8} mol/L is

a. acidic

b. basic

c. neutral

4. In an acidic solution,
   a. $[H^+]$ is equal to $[OH^-]$

   b. $[H^+]$ is greater than $[OH^-]$

   c. $[H^+]$ is less than $[OH^-]$

   d. $[H^+]$ are both $[OH^-] = 1 \times 10^{-7}$ mol/L

Learning Activity 5.5: Strong and Weak Acids and Bases

Use the Relative Strengths of Acids Table (Appendix D) found in the appendices section at the back of this course to complete these questions.

1. For each of the following:
   - Identify which of the two would behave as the acid and which would behave as the base.
   - Write balanced chemical equations for the reactions.
   - Indicate the acid/base conjugate acid base pairs.
   - Determine which direction would be favoured in each, using the K_a table.

   a. $H_3PO_4 + C_2H_3O_2^-$

   $H_3PO_4 + C_2H_3O_2^- \rightleftharpoons H_2PO_4^- + HC_2H_3O_2$

   acid, base, CB, CA

   The forward reaction is favoured, so H_3PO_4 is stronger than $HC_2H_3O_2$

   b. $C_2H_3O_2^- + HSO_4^-$

   $C_2H_3O_2^- + HSO_4^- \rightleftharpoons HC_2H_3O_2 + SO_4^{2-}$

   base, acid, CA, CB

   The forward reaction is favoured, so HSO_4^- is stronger than $HC_2H_3O_2$

   c. $SO_3^{2-} + NH_4^+$

   $SO_3^{2-} + NH_4^+ \rightleftharpoons HSO_3^- + NH_3$

   base, acid, CA, CB

   The reverse reaction is favoured, so HSO_3^- is the strongest acid

d. $HO_2^- + H_2CO_3$

$HO_2^- + H_2CO_3 \rightleftharpoons H_2O_2 + HCO_3^-$

base, acid, CA, CB

The forward reaction is favoured, so H_2CO_3 is the strongest acid.

2. Using the following ions:

- CO_3^{2-}, HCO_3^-, HPO_4^{2-}, CH_3COO^-, HSO_3^-

a. Identify which are acids and list in order of decreasing (strongest to weakest) strength.

Acids: Start at the top left and move down the left side.

HSO_3^-, HCO_3^-, HPO_4^{2-}

b. List those that behave as a base in order of decreasing strength.

Bases: Start from the lower right side of the strength chart. The lower the substance, the stronger the base.

CO_3^{2-}, HPO_4^{2-}, HCO_3^-, $C_2H_3O_2^-$, HSO_3^-

3. List the following substances in increasing order of strength (weakest to strongest) as acids and bases. Some may be used more than once.

- F^-, $H_2PO_4^-$, HSO_4^-, H_2S, $C_2H_3O_2^-$, HNO_2, HTe^-

Acids

HTe^- , $H_2PO_4^-$, H_2S , HNO_2, HSO_4^-

Bases

HSO_4^-, $H_2PO_4^-$, HTe^-, F^-, $C_2H_3O_2^-$

Learning Activity 5.6: Problem Solving with K_a and K_b

Solve the following problems. Be sure to show all your work and use correct units for your answers. If needed, use values from the K_a table that can be found in the appendices at the end of the course.

1. Calculate the pH of a 0.10 mol/L solution of hypochlorous acid, HOCl ($K_a = 3.5 \times 10^{-8}$)

To calculate pH, we need to calculate [H_3O^+]. Set up an ICE table.

	HOCl	$+ H_2O$	$\rightleftharpoons H_3O^+$	$+ OCl^-$
I	**0.10**		**0**	**0**
C	**– x**		**+ x**	**+ x**
E	**0.10 – x**		**x**	**x**

$$K_a = \frac{[H_3O^+][OCl^-]}{[HOCl]}$$

$$3.5 \times 10^{-8} = \frac{x^2}{0.10 - x}$$

$$3.5 \times 10^{-8} \approx \frac{x^2}{0.10}$$

$$\left(3.5 \times 10^{-8}\right)(0.10) = \left(\frac{x^2}{0.10}\right)(0.10)$$

$$\sqrt{3.5 \times 10^{-9}} = x$$

$$5.92 \times 10^{-5} = x$$

$$[H_3O^+] = x = 5.92 \times 10^{-5} \text{ mol/L}$$

$$pH = -\log[H_3O^+] = -\log\left(5.92 \times 10^{-5}\right) = 4.23$$

2. Calculate the K_a of a weak acid, HX, if a 0.25 mol/L solution has a pH of 4.40.

$HX + H_2O \rightarrow X^- + H_3O^+$

	HX	+ H_2O	$\rightleftharpoons$ X^-	+ H_3O^+
I	**0.25**		**0**	**0**
C	**– x**		**+ x**	**+ x**
E	**0.25 – x**		**x**	**x**

$$[H_3O^+] = 10^{-pH} = 10^{-4.40} = 3.98 \times 10^{-5} \text{ mol/L}$$

$$K_a = \frac{[H_3O^+][X^-]}{[HX]} = \frac{[x][x]}{0.25 - x} = \frac{x^2}{0.25 - x}$$

$$= \frac{\left(3.98 \times 10^{-5}\right)^2}{0.25 - x} = 6.0 \times 10^{-9}$$

3. Calculate the concentration of a solution of acetic acid ($K_a = 1.8 \times 10^{-5}$), which has a pH = 3.00.

$HC_2H_3O_2 + H_2O \rightarrow C_2H_3O_2^- + H_3O^+$

$$[H_3O^+] = 10^{-pH} = 10^{-3.00} = 1.0\times10^{-3}\ \text{mol/L}$$

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.8\times10^{-5} = \frac{(1.0\times10^{-3})^2}{[HC_2H_3O_2]}$$

$$[HC_2H_3O_2] = \frac{(1.0\times10^{-3})^2}{1.8\times10^{-5}}$$

$$[HC_2H_3O_2] = 0.056\ \text{mol/L}$$

Substitute values into mass action expression, then solve for $[HC_2H_3O_2]$.

4. Hydrazine (N_2H_4) is a weak base with a K_b of 3.0×10^{-6}. The reaction of hydrazine in water is $H_2NNH_2 + H_2O \rightleftharpoons H_2NNH_3^+ + OH^-$. Calculate the the pH of a 2.0 mol/L solution of hydrazine.

	H_2NNH_2	$+ H_2O$	$\rightleftharpoons OH^-$	$+ H_2NNH_3^+$
I	**2.0**		**0**	**0**
C	**– x**		**+ x**	**+ x**
E	**2.0 – x**		**x**	**x**

$$K_b = \frac{[H_2NNH_3^+][OH^-]}{[H_2NNH_2]}$$

$$3.0\times10^{-6} = \frac{x^2}{2.0-x}$$

$$3.0\times10^{-6} \approx \frac{x^2}{2.0}$$

$$(3.0\times10^{-6})(2.0) = \left(\frac{x^2}{2.0}\right)(2.0)$$

$$\sqrt{6.0\times10^{-6}} = x$$

$$2.45\times10^{-3} = x$$

$[OH^-] = 2.45 \times 10^{-3}$ mol/L

pOH = log $[OH^-]$ = log(2.45×10^{-3}) = 2.61

pH = 14 – pOH = 14 – 2.61 = 11.39

5. Calculate the percent dissociation of a 0.20 mol/L solution of the weak acid, HNO_2, if the pH of the solution is 4.20.

   $HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$

   If you find the hydronium ion concentration, you can calculate the percent dissociation.

   $[H_3O^+] = 10^{-pH} = 10^{-4.20} = 6.3 \times 10^{-5}$ **mol/L**

   $$\% \text{ dissociation} = \frac{\left[H_3O^+\right]}{\left[HNO_2\right]} \times 100\% = \frac{6.3 \times 10^{-5}}{0.20} \times 100\%$$

   % dissociation = 0.0315%

Learning Activity 5.7: Calculating an Unknown Concentration

1. Write the complete set of reactions that occur when the following acid and bases are reacted. Look back at Lesson 2 if you need some review first.

   a. Hydrochloric acid and sodium hydroxide

   Molecular: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

   Ionic: $H^+{}_{(aq)} + Cl^-{}_{(aq)} + Na^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)} + H_2O_{(l)}$

   Net Ionic: $H^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(l)}$

   b. Sulfuric acid and potassium hydroxide

   Molecular: $H_2SO_{4(aq)} + 2\,KOH_{(aq)} \rightarrow K_2SO_{4(aq)} + 2\,H_2O_{(l)}$

   Ionic: $2\,H^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\,K^+{}_{(aq)} + 2\,OH^-{}_{(aq)} \rightarrow 2\,K^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\,H_2O_{(l)}$

   Net Ionic: $H^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(l)}$

   c. Nitric acid and calcium hydroxide

   Molecular: $2\,HNO_{3(aq)} + Ca(OH)_{2(aq)} \rightarrow Ca(NO_3)_{2(aq)} + 2\,H_2O_{(l)}$

   Ionic: $2\,H^+{}_{(aq)} + NO_3^-{}_{(aq)} + Ca_2^+{}_{(aq)} + 2\,OH^-{}_{(aq)} \rightarrow Ca_2^+{}_{(aq)} + 2\,NO_3^-{}_{(aq)} + 2\,H_2O_{(l)}$

   Net Ionic: $H^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(l)}$

d. Phosphoric acid and lithium hydroxide

Molecular: $H_3PO_{4(aq)}$ + 3 $LiOH_{(aq)}$ → $Li_3PO_{4(aq)}$ + 3 $H_2O_{(l)}$

Ionic: 3 $H^+_{(aq)}$ + $PO_4^{3-}{}_{(aq)}$ + 3 $Li^+_{(aq)}$ + 3 $OH^-_{(aq)}$ → 3 $Li^+_{(aq)}$ + $PO_4^{3-}{}_{(aq)}$ + 3 $H_2O_{(l)}$

Net Ionic: $H^+_{(aq)}$ + $OH^-_{(aq)}$ → $H_2O_{(l)}$

e. Sulfuric acid and aluminum hydroxide

Molecular: 3 $H_2SO_{4(aq)}$ + 2 $Al(OH)_{3(aq)}$ → $Al_2(SO_4)_{3(aq)}$ + 6 $H_2O_{(l)}$

Ionic: 6 $H^+_{(aq)}$ + 3 $SO_4^{2-}{}_{(aq)}$ + 2 $Al^{3+}_{(aq)}$ + 6 $OH^-_{(aq)}$ → 2 $Al^{3+}_{(aq)}$ + 3 $SO_4^{2-}{}_{(aq)}$ + 6 $H_2O_{(l)}$

Net Ionic: $H^+_{(aq)}$ + $OH^-_{(aq)}$ → $H_2O_{(l)}$

2. Calculate the concentration of a solution of KOH if 140.0 mL of the base is needed to neutralize 22.5 mL of 0.175 mol/L acid HNO_3.

Write out the neutralization reaction.

2 $HNO_{3(aq)}$ + $KOH_{(aq)}$ → $KNO_{3(aq)}$ + $H_2O_{(l)}$

Calculate the number of moles of nitric acid.

moles HNO_3 = C × V = (0.175 mol/L)(0.0225 L) = 0.0039375 moles HNO_3

Since the molar ratio of HNO_3 to KOH is one-to-one, there are 0.0039375 moles of KOH in the sample.

$$\text{concentration} = \frac{\text{mol}}{\text{V}} = \frac{0.0039375\text{ moles}}{0.140\text{ L}}$$

The concentration of KOH solution is 0.0281 mol/L.

3. Determine the volume of 0.750 mol/L LiOH required to neutralize 56.0 mL of 0.350 mol/L H_2SO_4.

$H_2SO_{4(aq)}$ + 2 $LiOH_{(aq)}$ → $Li_2SO_{4(aq)}$ + 2 $H_2O_{(l)}$

Calculate the number of moles of sulphuric acid.

Moles of H_2SO_4 = C × V = (0.350 mol/L)(0.0560 L) = 0.0196 moles H_2SO_4

Use the molar ratio to determine the moles of LiOH present in the sample.

$$\text{moles LiOH} = 0.0196\text{ moles }H_2SO_4\left(\frac{2\text{ moles LiOH}}{1\text{ mole }H_2SO_4}\right) = 0.0392\text{ moles LiOH}$$

$$\text{Volume} = \frac{\text{mol}}{\text{C}} = \frac{0.0392\text{ moles}}{0.750\text{ mol/L}} = 0.05226\text{ L}$$

The volume of LiOH solution is 0.0523 L or 52.3 mL

4. Calculate the concentration of H_3PO_4 if 25.0 mL is required to neutralize 19.0 mL of 0.830 mol/L KOH.

$H_3PO_{4(aq)} + 3\ KOH_{(aq)} \rightarrow K_3PO_{4(aq)} + 3\ H_2O_{(l)}$

Calculate the number of moles of KOH.

Moles of KOH = C x V = (0.830 mol/L)(0.0190 L) = 0.01577 moles KOH

Use the molar ratio to determine the moles of H3PO4 present in the sample:

$$\text{moles } H_3PO_4 = 0.01577 \text{ moles KOH}\left(\frac{1 \text{ mole } H_3PO_4}{3 \text{ mole KOH}}\right) = 0.005257 \text{ moles } H_3PO_4$$

$$\text{Concentration} = \frac{\text{mol}}{\text{V}} = \frac{0.005257 \text{ moles}}{0.0250 \text{ L}} = 0.210 \text{ mol/L}$$

The concentration of the H_3PO_4 solution is 0.210 mol/L

5. What volume of 0.100 mol/L HCl is required to just neutralize 10.0 g of solid NaOH?

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Calculate the number of moles of NaOH.

$$\text{moles NaOH} = \frac{\text{mass}}{\text{Molar mass}} = \frac{10.0 \text{ g}}{40.0 \text{ g/mol}} = 0.250 \text{ moles NaOH}$$

Since the molar ratio is one-to-one, 0.250 moles of HCl is required to neutralize the NaOH.

$$\text{Volume} = \frac{\text{mol}}{\text{C}} = \frac{0.250 \text{ moles}}{0.100 \text{ mol/L}} = 2.50 \text{ L}$$

The volume of HCl solution needed is 2.50 L.

6. What volume of 0.250 mol/L nitric acid is needed to just neutralize 25.0 g of magnesium hydroxide?

$2\ HNO_{3(aq)} + Mg(OH)_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2\ H_2O_{(l)}$

Calculate the number of moles of $Mg(OH)_2$.

$$\text{moles } Mg(OH)_2 = \frac{\text{mass}}{\text{molar mass}} = \frac{25.0 \text{ g}}{58.3 \text{ g/mol}} = 0.4288 \text{ moles } Mg(OH)_2$$

Use the molar ratio to determine the moles of HNO_3 needed.

$$\text{moles } HNO_3 = 0.4288 \text{ moles } Mg(OH)_2 \left(\frac{2 \text{ moles } HNO_3}{1 \text{ mole } Mg(OH)_2}\right) = 0.8576 \text{ moles } HNO_3$$

$$\text{Volume} = \frac{\text{mol}}{C} = \frac{0.8576 \text{ moles}}{0.250 \text{ mol/L}} = 3.43 \text{ L}$$

The volume of HNO_3 solution needed is 3.43 L.

7. What mass of calcium hydroxide is needed to just neutralize 250. mL of 0.100 mol/L HCl? Assume the volume does not change.

   $2\,HCl_{(aq)} + Ca(OH)_{2(aq)} \rightarrow CaCl_{2(aq)} + 2\,H_2O_{(l)}$
   Calculate the number of moles of HCl.

   Moles of HCl = C × V = (0.100 mol/L)(0.250 L) = 0.0250 moles HCl

   Use the molar ratio to determine the moles of $Ca(OH)_2$ present in the solid sample.

$$\text{moles } Ca(OH)_2 = 0.0250 \text{ moles } HCl\left(\frac{1 \text{ mole } Ca(OH)_2}{2 \text{ mole } HCl}\right) = 0.0125 \text{ moles } Ca(OH)_2$$

   mass $Ca(OH)_2$ = mol × MM = (0.0125 moles)(74.1 g/mol) = 0.926 g

   The mass of solid, $Ca(OH)_2$, needed is 0.926 g.

Learning Activity 5.8: Predicting the Nature of Salts

Determine whether each of the following salts is acidic, basic, or neutral in water. For each answer, give a rationale to support your prediction, including the ions formed in each reaction.

1. Sodium nitrate

   $NaNO_3$

   Ions present: Na^+, NO_3^-

   Na^+: group 1 ion, therefore neutral

   NO_3^-: conjugate base of the strong acid HNO_3, therefore neutral

   $NaNO_3$ is neutral in water

2. Sodium sulphide

   Na_2S
   Ions present: Na^+, S^{2-}
   Na^+: group 1 ion, therefore neutral
   SO_4^{2-}: conjugate base of the weak acid HS^-, therefore basic
   $S^{2-} + H_2O \rightleftharpoons HS^- + OH^-$
   Na_2S is basic in water

3. Potassium chloride

   KCl
   Ions present: K^+, Cl^-
   K^+: group 1 ion, therefore neutral
   Cl^-: conjugate base of the strong acid HCl, therefore neutral
   KCl is neutral in water

4. Magnesium nitrate

   $Mg(NO_3)_2$
   Ions present: Mg^{2+}, NO_3^-
   Mg^{2+}: group 2 ion, therefore neutral
   NO_3^-: conjugate base of the strong acid HNO_3, therefore neutral
   $Mg(NO_3)^2$ is neutral in water

5. Sodium carbonate

   Na_2CO_3
   Ions present: Na^+, CO_3^{2-}
   Na^+: group 1 ion, therefore neutral
   CO_3^{2-}: conjugate base of the weak acid HCO_3^-, therefore basic
   $CO_32^- + H_2O \rightleftharpoons HCO_3^- + OH^-$
   Na_2CO_3 is basic in water

6. Ammonium iodide

   NH_4I
   Ions present: NH_4^+, I^-
   NH_4^+: conjugate acid of weak base NH_3, therefore acidic
   $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$
   I^-: conjugate base of the strong acid HI, therefore neutral
   NH_4I is acidic in water

7. Potassium acetate

   $KC_2H_3O_2$
   Ions present: K^+, $C_2H_3O_2^-$
   K^+: group 1 ion, therefore neutral
   $C_2H_3O_2^-$: conjugate base of the weak acid $HC_2H_3O_2$, therefore basic
   $C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$
   $KC_2H_3O_2$ is basic in water

8. Calcium oxide

   CaO
   Ions present: Ca^{2+}, O^{2-}
   Ca^{2+}: group 1 ion, therefore neutral
   O^{2-}: oxides hydrolyze water, therefore basic
   $O^{2-} + H_2O \rightleftharpoons OH^- + OH^-$
   CaO is basic in water

Notes

GRADE 12 CHEMISTRY (40S)

Module 6
Electrochemistry

MODULE 6:
ELECTROCHEMISTRY

Introduction to Module 6

As the reliance on hydrocarbons for sources of energy becomes less, the need for alternative energy sources will become greater. Electrochemistry will play a large role in the search for alternative energy sources.

Many chemical reactions between metals and aqueous solutions will produce energy. This is known as an electrochemical cell, and it allows us to make batteries. You will have an opportunity, in this unit, to construct a battery made of lemons, copper pennies, and a galvanized nail. This will allow you to see a chemical reaction produce electricity.

There are other applications in the real world in which electrochemistry plays a part. Precious metals like gold and silver are often plated onto less expensive metals for protection from oxidation.

Note

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.

Assignments in Module 6

Note

Please remember that Assignment 1.9: Researching a Redox Application is to be submitted with the assignments in this module. The marking rubric is found in Module 1 and at the end of this module.

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	Predicting Spontaneity
2	Assignment 6.2	History of the Voltaic Cell
3	Assignment 6.3	The Functioning Voltaic Cell
4	Assignment 6.4	Calculating Standard Reduction Potentials
5	Assignment 6.5	Electroplating
6	Assignment 6.6	Using Faraday's Law
Module 1	Assignment 1.9	Researching Redox Reactions

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 4 to 6, and is worth 30 percent of your final mark in the course. To do well on the final examination, you should review all the work you complete in Modules 4, 5, and 6, including all the learning activities and assignments. You will write the final examination under supervision.

LESSON 1: ACTIVITY SERIES (2 HOURS)

Lesson Focus

SLO C12-6-01: Develop an activity series experimentally.

SLO C12-6-02: Predict the spontaneity of reactions using an activity series.

Lesson Introduction

Hydrochloric acid (HCl) will react with zinc, but not with copper or gold. Gold is used in jewellery because of its high lustre, its malleability, and its resistance to oxidation. The reaction of zinc with HCl is a spontaneous reaction, meaning that it occurs without any added energy. The corrosion of iron to form iron oxide is also a spontaneous reaction.

In this lesson, you will learn how to determine which substances in a chemical group are the easiest to oxidize and which are the easiest to reduce.

Spontaneous Reactions

A **spontaneous reaction** is one that occurs without any added energy. For example, when you add solid zinc to hydrochloric acid, fizzing occurs, followed by the loss of mass of the zinc. Eventually, all the solid zinc disappears. The reaction occurs as follows:

$$Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

The reaction above is a **single replacement reaction**, where the zinc atoms replace the hydrogen atoms. Zinc is oxidized and hydrogen is reduced in this reaction. All single replacement reactions are redox reactions.

For this reaction to occur, there must be a competition for zinc's electrons. If this reaction is to be spontaneous, hydrogen must be a strong enough oxidizing agent to remove the electrons from zinc, and zinc must have an **electron affinity** that is low enough for hydrogen to remove its electrons.

Not all reactions are spontaneous. If you place copper into hydrochloric acid, no reaction occurs. This is not a spontaneous reaction because hydrogen ions are not strong enough, as oxidizing agents, to remove copper's electrons.

Oxidation and Reduction Review

When two substances react, the **strongest oxidizing agent will accept electrons from another substance, becoming reduced in the process**. The **strongest reducing agent will lose its electrons, becoming oxidized**. If you need to review these concepts, revisit Module 1. It is important to note that metals do not act as oxidizing agents. In other words, they will not gain electrons to have negative oxidation numbers.

Note

- When an atom goes from a 0 charge to a positive charge, electrons have been removed.
- When an atom or ion becomes more positively charged, oxidation has taken place.

Here is an example that you saw in Module 1. In the following reaction, iron slowly turns to rust.

$$2\,Fe_{(s)} + 3\,O_{2(g)} \rightarrow 2\,Fe_2O_{3(s)}$$

Written in the ionic form, the reaction becomes

$$2\,Fe^0 + 3\,O_2{}^0 \rightarrow 2\,Fe_2{}^{3+}O_3{}^{2-}$$

In this example, iron oxidizes (gains oxygen) to form iron oxide. Since iron went from a charge of 0 to a charge of 3^+, you know that three electrons were lost (or removed). Iron, having lost three electrons, now has three **more** protons than electrons, resulting in a charge of 3^+.

To help you associate oxidation with removal of electrons, you were introduced to the acronym LEO. Each letter of the acronym represents a word that creates the reminder “**L**osing **E**lectrons is **O**xidation.”

Competition for Electrons

Consider the reaction between copper and silver as an example. When copper metal is placed into a solution of silver ions, crystals of silver metal are produced on the copper strip. The solution begins to turn from colourless to blue, indicating that copper metal is converted into copper ions.

The following equations describe the results of this reaction:

$$2\ Ag^{+}_{(aq)} + Cu^{0}_{(s)} \rightarrow 2\ Ag^{0}_{(s)} + Cu^{2+}_{(aq)}$$

$$Ag^{0}_{(s)} + Cu^{2+}_{(aq)} \rightarrow \text{no reaction}$$

Note that the second reaction is the reverse of the first. Based on this information, you can conclude the following:

- Ag^{+} ions gain electrons and are, therefore, reduced.
- Ag^{+} ions can oxidize Cu metal.
- Cu^{0} loses electrons and is, therefore, oxidized.
- Cu^{2+} cannot oxidize Ag metal.
- Cu metal can reduce Ag^{+} ions.
- Ag metal cannot reduce Cu^{2+} ions.

If this was a tug-of-war between silver ions and copper ions over the electrons, the silver ions would win. As such, silver ions are a stronger oxidizing agent than copper ions, and copper metal is a stronger reducing agent than silver metal. Silver ions have the highest affinity for electrons.

What is an Activity Series?

If you did many experiments, like in the simulations in the next section, you could set up an **activity series** that lists substances in order of their ability as oxidizing and reducing agents. A true activity series, called the **Activity Series of Metals**, will only list metals becoming an ion. This is an invaluable aid to predicting the products of replacement reactions, and it can also be used as an aid in predicting products of some other reactions. The following summary points will help you make better use of the activity series:

- Each element on the list from a compound replaces any of the elements below it. The larger the interval between elements, the more vigorous the reaction.
- The first five elements (lithium to sodium) are known as very active metals and they react with cold water to produce the metal hydroxide and hydrogen gas.

- The next four metals (magnesium to chromium) are considered active metals and they will react with very hot water or steam to form the oxide and hydrogen gas.
- The oxides of all of these first metals resist reduction by H_2.
- The next five metals (iron to lead) replace hydrogen in HCl and diluted sulfuric and nitric acids. Their oxides undergo reduction by heating with H_2, carbon, and carbon monoxide.
- The metals lithium to copper can combine directly with oxygen to form the oxide.
- The last three metals (mercury to gold) are often found free in nature; their oxides decompose with mild heating and they form oxides only indirectly.

Although most metals are electropositive in nature and lose electrons in a chemical reaction, they do not react with the same vigour or speed. The greater the ease with which an element loses its electrons and acquires a positive charge, the greater is its reactivity. Further, the greater the number of shells and the lesser the number of valence electrons, the greater is the reactivity of the metal. The activity series of metals arranges all metals in order of their decreasing chemical activity. As we go down the activity series from lithium to gold, the ease with which a metal loses electrons and forms positive ions in solutions decreases.

The most active metal, lithium, is at the top of the list; and the least reactive metal, gold, is at the bottom of the list. Although hydrogen is a non-metal, it is included in the activity series due to the fact that it behaves like a metal in most chemical reactions.

Activity Series of Metals in Aqueous Solutions

Description	Metal	Chemical Symbol	
■ These metals displace hydrogen from water. ■ These elements are very reactive and react readily to form compounds.	Lithium	Li	↑ Increase in Ease of Oxidation ↓ Increase in Ease of Reduction
	Potassium	K	
	Barium	Ba	
	Calcium	Ca	
	Sodium	Na	
■ These metals displace hydrogen from acids.	Magnesium	Mg	
	Aluminum	Al	
	Zinc	Zn	
	Chromium	Cr	
	Iron	Fe	
	Cadmium	Cd	
	Nickel	Ni	
	Tin	Sn	
	Lead	Pb	
	Hydrogen	H	
■ These elements are more stable and form compounds less readily. ■ The metals do not displace hydrogen from acids or water.	Copper	Cu	
	Silver	Ag	
	Mercury	Hg	
	Platinum	Pt	
	Gold	Au	

An activity series can also list other species in order of reactivity. Most North American activity series are listed as standard reduction potentials at 25°C and ionic concentrations of 1 mol/L solution.

Many activity series will be ranked in terms of reduction reactions, like the one that follows. The better **oxidizing agents**, like permanganate and peroxide, are found at the bottom left of the table. They will draw electrons away from almost any substance to become reduced.

↑ Decreasing Ability as an Oxidizing Agent	$Li^+ + e^-$	→	$Li_{(s)}$	Strongest Reducing Agents Easiest to Oxidize
	$Rb^+ + e^-$	→	$Rb_{(s)}$	
	$K^+ + e^-$	→	$K_{(s)}$	
	$2\ H_2O + 2e^-$	→	$2\ OH^- + H_{2(g)}$	
	$Zn^{2+} + 2e^-$	→	$Zn_{(s)}$	
	$Cr^{3+} + 3e^-$	→	$Cr_{(s)}$	Decreasing Ability as a Reducing Agent ↓
	$Sn^{2+} + 2e^-$	→	$Sn_{(s)}$	
	$Pb^{2+} + 2e^-$	→	$Pb_{(s)}$	
	$Fe^{3+} + 3e^-$	→	$Fe_{(s)}$	
	$2\ H^+ + 2e^-$	→	$H_{2(g)}$	
	$Cu^{2+} + 2e^-$	→	$Cu_{(s)}$	
	$Cu^+ + e^-$	→	$Cu_{(s)}$	
	$I_{2(s)} + 2e^-$	→	$2\ I^-$	
	$O_{2(g)} + 2H^+ + e^-$	→	H_2O_2	
	$Li^+ + e^-$	→	$Li_{(s)}$	
	$Fe^{3+} + e^-$	→	Fe^{2+}	
	$NO_3^-+ 2\ H^+ + e^-$	→	$NO_{2(g)} + H_2O$	
	$Ag^+ + e^-$	→	$Ag_{(s)}$	
	$NO_3^-+ 4\ H^+ + 3e^-$	→	$NO_{(g)} + 2\ H_2O$	
	$Br_{2(g)} + 2e^-$	→	$2\ Br^-$	
Strongest Oxidizing Agents Easiest to Reduce	$½O_{2(g)} + 2\ H^+ + 2e^-$	→	H_2O	
	$Au^{3+} + 3e^-$	→	$Au_{(s)}$	
	$MnO_4^- + 8\ H^+ + 5e^-$	→	$Mn^{2+} + 4\ H_2O$	
	$H_2O_2 + 2\ H^+ + 2e^-$	→	$2\ H_2O$	
	$F_{2(g)} + 2e^-$	→	$2\ F^-$	

Lithium and potassium ions are located at the top left-hand side of the table. They do not readily accept electrons. In fact, they are more than happy to remain as ions.

Substances such as lithium and potassium, which are easiest to oxidize (best **reducing agents**), are found at the top right-hand side of the table. These elements are not found as native or pure elements in nature and will even react violently with water. They are quite eager to give their electrons to any substance in order to become oxidized.

Use the following simple examples to test your understanding of the activity series:

1. What species is the strongest oxidizing agent?

   Answer: $F_{2(g)}$ (bottom, left-hand side of the table)
2. What species is the least easily reduced?

   Answer: $Li^{+}_{(aq)}$ (top, left-hand side of the table)
3. What species is the weakest reducing agent?

   Answer: $F^{-}_{(aq)}$ (bottom, right-hand side of the table)

Metals like gold are found close to the bottom, on the right-hand side of the table. Gold does not oxidize easily and is often found in its elemental form in nature.

Note

The stronger the oxidizing agent, the weaker the reducing ability.

The stronger the reducing agent, the weaker the oxidizing ability.

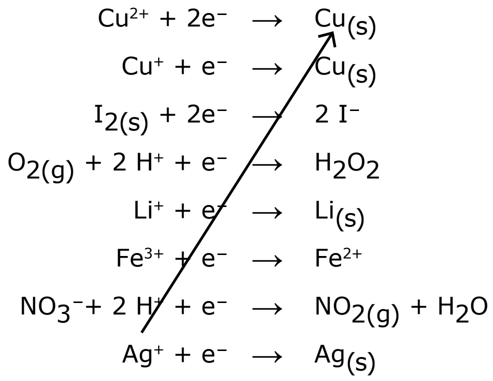
Reactions of Metals and Metal Ions Experiment (Simulation)

Use a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/ to find simulations of experiments where different metals are placed in various aqueous solutions.

The Spontaneity Rule

Based on the results of experiments with chemical reactions, chemists discovered that certain species reacted spontaneously, whereas others did not react. You may recall from Grade 10 Science that alkali metals react to become positive ions as they lose electrons to become stable and end up having a similar electronic structure to the inert gases.

The activity series can be used to predict if a reaction will be spontaneous. The **spontaneity rule** states that any species on the left-hand side of the activity series **(written as reduction reactions)** will react spontaneously and oxidize any species on the **right-hand side which is above it**. In other words, spontaneous reactions are "up to the right." Following is the example of copper and silver, if you use the activity series:



$$\begin{aligned}
Cu^{2+} + 2e^- &\longrightarrow Cu_{(s)} \\
Cu^{+} + e^- &\longrightarrow Cu_{(s)} \\
I_{2(s)} + 2e^- &\longrightarrow 2\ I^- \\
O_{2(g)} + 2\ H^+ + e^- &\longrightarrow H_2O_2 \\
Li^{+} + e^- &\longrightarrow Li_{(s)} \\
Fe^{3+} + e^- &\longrightarrow Fe^{2+} \\
NO_3^- + 2\ H^+ + e^- &\longrightarrow NO_{2(g)} + H_2O \\
Ag^{+} + e^- &\longrightarrow Ag_{(s)}
\end{aligned}$$

Copper metal is "up and to the right" of silver ions, so the reaction is spontaneous. Silver metal is **not** "up and to the right" of copper ions, so the reaction is not spontaneous. Silver ions will react with any metal that is above it in the activity series on page 10 of this module (for example, lead, tin, magnesium, etc.).

Example 1

Will tin strips in hydrochloric acid react?

Step 1: Write the formulas for the reactants.

There is tin metal and hydrochloric acid. For the acid, the reactive component is the hydrogen ion. The reactants are

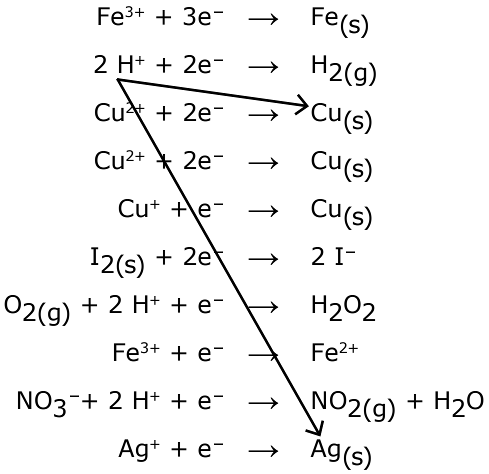
$$Sn_{(s)} + H^{+}_{(aq)} \rightarrow$$

Step 2: Check the activity series.

$Cr^{3+} + 3e^{-} \longrightarrow Cr_{(s)}$
$Sn^{2+} + 2e^{-} \longrightarrow \boxed{Sn_{(s)}}$
$Pb^{2+} + 2e^{-} \longrightarrow Pb_{(s)}$
$Fe^{3+} + 3e^{-} \longrightarrow Fe_{(s)}$
$2\,H^{+} + 2e^{-} \longrightarrow H_{2(g)}$
$Cu^{2+} + 2e^{-} \longrightarrow Cu_{(s)}$
$Cu^{2+} + 2e^{-} \longrightarrow Cu_{(s)}$

Tin is "up and to the right" of hydrogen ions; therefore, this is a spontaneous reaction.

Notice that copper and silver metal are "down and to the right" of hydrogen ions. This means they will not react with hydrochloric acid.



$Fe^{3+} + 3e^{-} \longrightarrow Fe_{(s)}$
$2\,H^{+} + 2e^{-} \longrightarrow H_{2(g)}$
$Cu^{2+} + 2e^{-} \longrightarrow Cu_{(s)}$
$Cu^{2+} + 2e^{-} \longrightarrow Cu_{(s)}$
$Cu^{+} + e^{-} \longrightarrow Cu_{(s)}$
$I_{2(s)} + 2e^{-} \longrightarrow 2\,I^{-}$
$O_{2(g)} + 2\,H^{+} + e^{-} \longrightarrow H_2O_2$
$Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$
$NO_3^{-} + 2\,H^{+} + e^{-} \longrightarrow NO_{2(g)} + H_2O$
$Ag^{+} + e^{-} \longrightarrow Ag_{(s)}$

Other metals such as lead, iron, zinc, and so forth will react with hydrochloric acid.

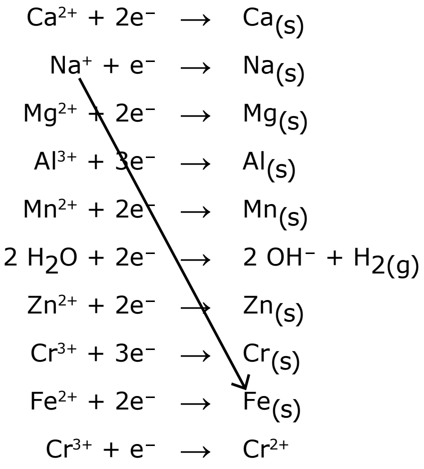
Example 2

Will a solution of salt (sodium chloride) react with a steel container?

Step 1: Steel contains iron and salt contains sodium ions, so the reaction is

$$Fe_{(s)} + Na^{+}_{(aq)} \rightarrow$$

Step 2: Check the activity series.



$$\begin{aligned}
Ca^{2+} + 2e^{-} &\rightarrow Ca_{(s)} \\
Na^{+} + e^{-} &\rightarrow Na_{(s)} \\
Mg^{2+} + 2e^{-} &\rightarrow Mg_{(s)} \\
Al^{3+} + 3e^{-} &\rightarrow Al_{(s)} \\
Mn^{2+} + 2e^{-} &\rightarrow Mn_{(s)} \\
2\,H_2O + 2e^{-} &\rightarrow 2\,OH^{-} + H_{2(g)} \\
Zn^{2+} + 2e^{-} &\rightarrow Zn_{(s)} \\
Cr^{3+} + 3e^{-} &\rightarrow Cr_{(s)} \\
Fe^{2+} + 2e^{-} &\rightarrow Fe_{(s)} \\
Cr^{3+} + e^{-} &\rightarrow Cr^{2+}
\end{aligned}$$

Iron is "down and to the right" of sodium ions; therefore, the reaction is not spontaneous.

$$Fe_{(s)} + Na^{+}_{(aq)} \rightarrow \text{no reaction}$$

Predicting an Activity Series

If you are given experimental data, such as that from the copper and silver experiment, you can construct your own activity series.

Example 3

Given the following experimental data, where we observed how several solids reacted with solutions of Cu^{2+} and Co^{2+}, arrange the following in increasing order of oxidizing ability.

1. $Co^{2+} + In_{(s)} \rightarrow Co_{(s)} + In^{2+}$
2. $Cu^{2+} + Co_{(s)} \rightarrow Cu_{(s)} + Co^{2+}$
3. $Cu^{2+} + Pd_{(s)} \rightarrow$ no reaction

Step 1: Arrange the **reduction reactions** from weakest oxidizing agent to strongest.

Work with one equation/reaction at a time and determine the oxidizing agent.

1. $Co^{2+} + In_{(s)} \rightarrow Co_{(s)} + In^{2+}$

In has electrons at the start of the reaction and is then oxidized. Co^{2+} has a greater affinity for electrons and is reduced. Co^{2+} is, therefore, the oxidizing agent.

2. $Cu^{2+} + Co_{(s)} \rightarrow Cu_{(s)} + Co^{2+}$

Co has electrons at the start of the reaction and is then oxidized. Cu^{2+} has a greater affinity for electrons and is reduced. Cu^{2+} is, therefore, the oxidizing agent.

3. $Cu^{2+} + Pd_{(s)} \rightarrow$ no reaction

   Cu^{2+} is unable to oxidize Pd.

Step 2: Arrange in columns according to the spontaneity rule.

According to equation 1, $In_{(s)}$ must be "up and to the right" of Co^{2+}. Arrange the two reduction equations accordingly.

$$In^{2+} + 2e^- \rightarrow In_{(s)}$$
$$Co^{2+} + 2e^- \rightarrow Co_{(s)}$$

According to the second equation, $Co_{(s)}$ must be "up and to the right" of Cu^{2+}.

$$In^{2+} + 2e^- \rightarrow In_{(s)}$$
$$Co^{2+} + 2e^- \rightarrow Co_{(s)}$$
$$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$$

According to the third equation, $Pd_{(s)}$ must be below Cu^{2+}.

$$In^{2+} + 2e^- \rightarrow In_{(s)}$$
$$Co^{2+} + 2e^- \rightarrow Co_{(s)}$$
$$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$$
$$Pd^{2+} + 2e^- \rightarrow Pd_{(s)}$$

From the experimental data, Pd^{2+} is the strongest oxidizing agent of the group.

Example 4

A zinc metal strip is placed into a 1.0 mol/L solution of copper (II) nitrate. Is the reaction spontaneous or not?

Step 1: Determine the species available for reacting.

In this example, they are $Zn^0_{(s)}$, $Cu^{2+}_{(aq)}$, and $NO_3^-{}_{(aq)}$.

$$Zn^0_{(s)} + Cu^{2+}_{(aq)} + NO_3^-{}_{(aq)} \rightarrow ?$$

Step 2: Determine which species is oxidized and which is reduced.

$NO_3^-{}_{(aq)}$ ions will not react unless the solution is acidic.

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu^0_{(s)}$$

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn^0_{(s)}$$

According to the table on page 10 of this module, Cu^{2+} has a greater affinity for electrons than Zn^{2+} and so Cu^{2+} will attract electrons to become reduced in the process to Cu^0. This causes Zn^0 to give up electrons and become oxidized to Zn^{2+}.

Step 3: Determine the net reaction as well as whether it is spontaneous or not.

The two reactions are as follows:

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu^0_{(s)}$$

$$\underline{Zn^0_{(s)} \rightarrow 2e^- + Zn^{2+}_{(aq)} \qquad\qquad\qquad}$$

$$Cu^{2+}_{(aq)} + Zn^0_{(s)} \rightarrow Cu^0_{(s)} + Zn^{2+}_{(aq)}$$

The reaction is spontaneous.

Learning Activity 6.1: Using the Activity Series

Use the activity series table to complete the following questions:

1. Indicate whether or not the following reactions will occur based on their relative position on the activity series table. If the reaction occurs, complete the equation.

   a. $Cu^+ + Pb \rightarrow$

   b. $Sn + Cr^{3+} \rightarrow$

   c. $NO_3^- + 4H^+ + Fe^{2+} \rightarrow$

   d. $Pb^{2+} + Ag_2S \rightarrow$

2. Answer the following questions about these species: Au^{3+}, $Cr^0_{(s)}$, Sn^{2+}, and Br^-.

   a. Which is the most easily reduced or, in other words, the strongest oxidizing agent?
   b. Which has the greatest affinity for electrons?
   c. Which is the least easily oxidized?
   d. Which is the most easily oxidized?
   e. Which will oxidize Sn^{2+} to Sn^{4+}?
   f. Which will reduce $F_{2(g)}$ to $2F^-_{(aq)}$?

3. Consider the following equations:

   $K + LiCl \rightarrow ?$

   $Li + KCl \rightarrow ?$

   Why can't they both occur spontaneously?

4. How can you determine which atom will have a stronger pull on, or "win the competition for," another atom? Provide an example.

Check the answer key.

Lesson Summary

In this lesson, you learned about the activity series that lists substances in order of their ability as oxidizing and reducing agents. You then used the activity series to predict whether a reaction will be spontaneous. You were introduced to the spontaneity rule, which states that any species on the left-hand side of the activity series, written as reductions, will oxidize any species on the right-hand side that is above it.

Assignment 6.1: Predicting Spontaneity (13 marks)

1. In each of the following, indicate whether the reaction is spontaneous or not. *(1 mark each)*

   a. A copper strip of metal is placed in a solution of zinc nitrate.

   b. A lead strip is placed in a solution of zinc nitrate.

   c. Hydrochloric acid is placed in a container of lead metal.

   d. Hydrochloric acid is placed in a container of zinc metal.

   e. A copper strip is placed in a solution of lead (II) nitrate.

   f. A zinc strip is placed in a solution of copper (II) nitrate.

   g. A solution of hydrochloric acid is placed in a container of copper.

   h. Lead (II) nitrate solution is placed in a container of zinc metal.

   i. A lead strip is placed in a solution of copper (II) nitrate.

   j. Fluorine gas is bubbled in a sodium bromide solution.

continued

Assignment 6.1: Predicting Spontaneity (continued)

2. For each pair of metals listed below, indicate which one is more easily oxidized. *(0.5 mark x 6 = 3 marks)*

   a. Hg, Cu

   ____

   b. Ca, Al

   ____

   c. Ni, Mg

   ____

   d. Sn, Ag

   ____

   e. Pb, Zn

   ____

   f. Cu, Al

   ____

LESSON 2: HISTORY OF THE VOLTAIC CELL (2 HOURS)

Lesson Focus

SLO C12-6-03: Outline the historical development of voltaic (galvanic) cells.
Include: contributions of Luigi Galvani and Alessandro Volta

Lesson Introduction

Chances are that you use a battery in some form every day. The battery in your car generates an electrical current so you can start the engine. The battery in your portable music player and cell phone also generates an electric current. All batteries convert chemical potential energy into electrical energy by using redox reactions. In this lesson, you will be introduced to contributions made by early chemists to electrochemistry.

Development of the Voltaic Cell (Galvanic Cell)

Luigi Galvani (1737–1798) discovered that muscle contractions result from electrical stimulation. He believed that electricity was a natural entity produced only in animals because he observed that electrical charges could make frog legs jump, even if the legs were no longer attached to a frog. While cutting a frog leg, Galvani's steel scalpel touched a brass hook that was holding the leg in place and the leg twitched. Further experiments confirmed this effect, and Galvani was convinced that he was seeing the effects of what he called "animal electricity" (the life force within the muscles of the frog).

Alessandro Giuseppe Antonio Anastasio Volta (1745–1824), an Italian physicist, studied Arrhenius' work on electrolytes. Volta studied under Galvani. Volta didn't think that electricity was exclusive to animals so he set out to disprove Galvani's hypothesis. In 1799, after extensive experimentation, Volta took metal disks and piled them one on top of the other, separated by strips of leather soaked in a salt or acid solution. He discovered that this "pile" generated an electric current. This was the first battery ever created!

Volta would go on to test many different combinations of metals and salt solutions, increasing the number of metal disks used. He would test the strength of the current produced by touching the ends of the battery to his face (don't try this at home). He eventually ended this testing method after one of the tests left his ears ringing.

Napoleon was so impressed by Volta's invention he made Volta a count and a senator.

In order to learn more about Alessandro Volta, use a search engine such as www.google.ca/.

Electrical Terms

Before discussing how electricity is produced in a voltaic cell, you must have a general understanding of some basic electrical terms. You may know some of these terms from, among other courses, Grade 9 Science.

The term **electric current** refers to a flow of electrons. This might seem familiar since a current in water refers to moving water. An electrical current will flow through any conductor, including electrolyte solutions. Electrons flow in and out of an electrochemical cell through conductors—usually metals or graphite—called electrodes.

The battery commonly used in cameras and flashlights is 1.5 volts, and a car battery is 12 volts. **Voltage** (V) is a unit of cell potential or **electrical potential** ($E°$). Cell or electrical potential is the ability of an electrochemical cell to do electrical work, or its ability (force) to move electrons from one electrode to another. Using the water analogy, cell potential is similar to water pressure: the greater the pressure, the higher the voltage and the more work that can be done.

Components of Electrochemical Cells

According to the law of conservation of energy, energy cannot be created or destroyed; it is simply converted from one form of energy to another. In an electrochemical cell, the chemical energy from a spontaneous redox reaction is converted to electrical energy.

Volta was able to generate an electrical current because he was able to take advantage of a spontaneous redox reaction. He separated the oxidation and reduction reactions and allowed the transfer of electrons to occur through a wire.

Note

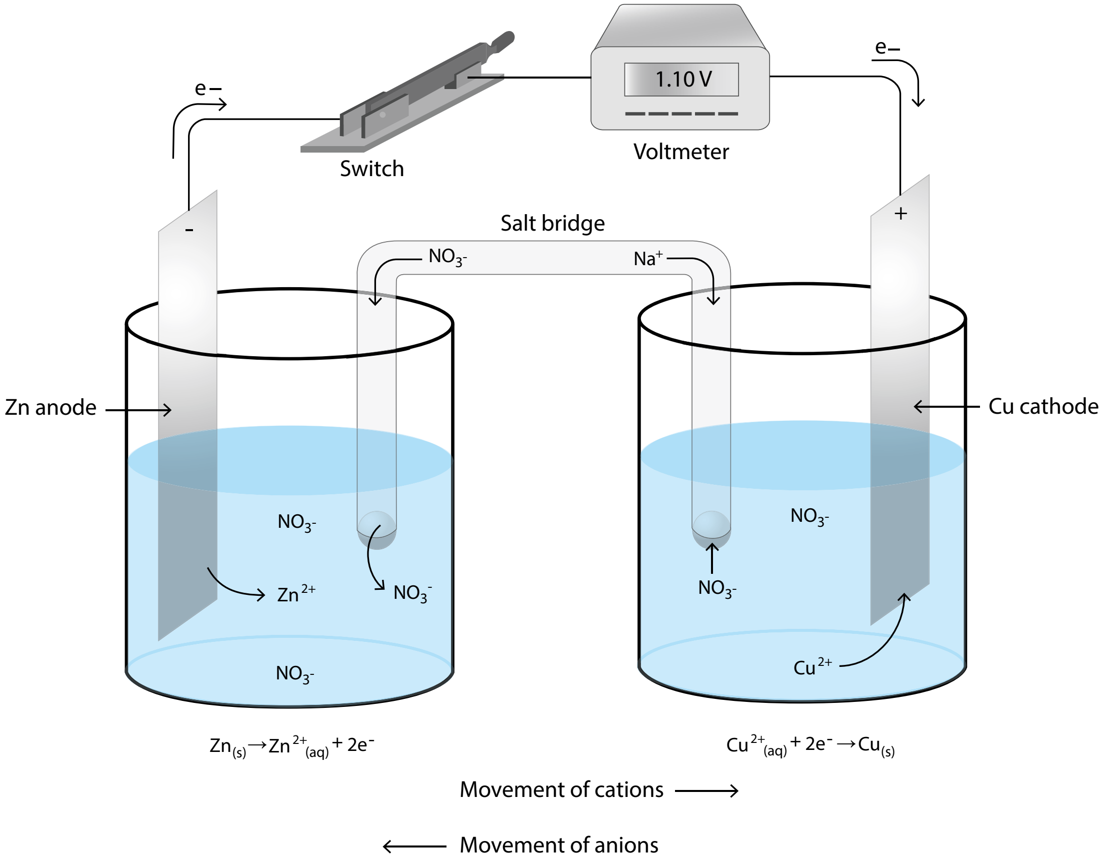
An **electrochemical cell** converts chemical energy into electrical energy from a spontaneous redox reaction.

The diagram below shows a typical electrochemical apparatus in a chemistry lab.

The oxidation reaction happens in one container and the reduction reaction happens in another container. Each container is called a **half-cell**. The **electrodes** are connected by a wire, and the electrons are released by the oxidation half-cell and transferred to the electrode in the reduction half-cell through the wire. If a light bulb is connected to the circuit and the voltage is high enough, it can illuminate. The following is an example of an electrochemical cell made from the spontaneous reaction:

$$Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$$

In this cell, zinc metal will be reduced and copper ions will be oxidized. The cell would look like the diagram below.



Note

The **anode** is the electrode where **oxidation** occurs. Electrons are **lost** from this electrode. It is normally labelled the **negative** electrode.

The **cathode** is the electrode where **reduction** occurs. Electrons are **gained** by positive ions at the surface of this electrode. It is normally labelled the **positive** electrode.

A common phrase used to remember which electrode is which is “A **red cat** ate **an ox**.”

REDuction = **CAT**hode; **AN**ode = **OX**idation

Learning Activity 6.2: History of the Voltaic Cell

1. You have been introduced to a short history of the voltaic cell in Lesson 6.2. Construct a timeline to outline some of the major events in electrochemistry. Be sure to include Luigi Galvani’s discovery of “animal electricity,” Volta’s first battery, Gaston Plante’s lead battery, and, finally, Waldmar Jungner’s Ni-Cad battery. More information on the contribution of the four individuals to the voltaic cell can be found by conducting research.

Check the answer key.

Lesson Summary

In this lesson, you learned that a chemical reaction can be used to create an electric current. You were introduced to electrochemical cells as an apparatus that uses a spontaneous redox reaction to produce electrical energy. In the next lesson, you will further your study of electrochemical cells at the visual, particulate, and symbolic levels.

Assignment 6.2: History of the Voltaic Cell (12 marks)

Conduct a short research to investigate the contributions of other scientists to our knowledge of electrochemical cells. Provide a brief overview of two such people and their work. The following sources show some of the history of the electrochemical cell together with images of some of the original voltaic cells:

- http://physics.kenyon.edu/EarlyApparatus/Electricity/Electrochemical_Cell/Electrochemical_Cell.html
- www.corrosion-doctors.org/Biographies/GalvaniBio.htm
- www.energyquest.ca.gov/scientists/galvani.html

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

If you do not have access to the Internet, please contact your tutor/marker or access your local library.

The example below has been provided for you to follow.

Name of scientist *(1 mark)*

Georges Leclanche

Electrodes and electrolyte used *(3 marks)*

carbon, zinc, sal ammoniac

Uses(s) of this electrochemical cell *(1 mark)*

intermittent ringing bells

Source of information *(1 mark)*

http://physics.kenyon.edu/EarlyApparatus/Electricity/Electrochemical_Cell/Electrochemical_Cell.html

continued

Assignment 6.2: History of the Voltaic Cell (continued)

1. Name of scientist: *(1 mark)*

   Electrodes and electrolyte used: *(3 marks)*

   Use(s) of this electrochemical cell: *(1 mark)*

   Source of information: *(1 mark)*

2. Name of scientist: *(1 mark)*

   Electrodes and electrolyte used: *(3 marks)*

   Use(s) of this electrochemical cell: *(1 mark)*

   Source of information: *(1 mark)*

LESSON 3: HOW AN ELECTROCHEMICAL CELL WORKS (2.5 HOURS)

Lesson Focus

SLO C12-6-04: Explain the operation of a voltaic (galvanic) cell at the visual, particulate, and symbolic levels.
Include: writing half-cell reactions, the overall reaction, and short-hand (line) notation

SLO C12-6-05: Construct a functioning voltaic (galvanic) cell and measure its potential.

Lesson Introduction

In the previous lesson, you learned about the contributions of Alessandro Volta and Luigi Galvani. Now that you are familiar with the electrochemical cell, you will learn how to write half-cell reactions and the overall reaction for a functioning cell.

Building an Electrochemical Cell

Note

Voltaic, galvanic, and **electrochemical** cells are all synonymous terms. For the purposes of this course, the term *electrochemical* will be used more often than others.

There are many chemical and electrical processes that occur in an electrochemical cell. So far in this module, you have learned that oxidation and reduction reactions take place in an electrochemical cell. Redox reactions take place in electrochemical cells without the help of an external current, and are, therefore, known as spontaneous.

An electrochemical cell consists of two half-reactions connected by an external wire that will carry the electric current. Remember also that electric current is the flow of electrons. The circuit must be complete for the cell to operate.

There are generally two metals in the cell that serve as electrodes. The electrode where oxidation takes place is called the **anode**. The metal undergoing oxidation will lose electrons, which will pass through the external wire to the other electrode. The electrode where reduction takes place is called the **cathode**.

The force "driving" the electrons is called the EMF, or **electromotive force**, and is measured in volts. You will work with EMF (symbol E^o) in the next lesson.

The Salt Bridge

The movement of ions is essential to the operation of the cell. At the anode, oxidation forms soluble positive ions. If these ions were allowed to accumulate at this electrode, the solution would become so positive that electrons would soon stop flowing to the cathode because of electrostatic attraction. At the cathode, soluble positive ions are reduced forming a solid. This removal of positive ions from the solution would eventually make the area around the cathode so negative that electrons coming from the anode would be repelled. To counter the accumulation of ions around the electrodes, there must be an internal connection between the cells that allows the movement of ions. This is called the **salt bridge.** In high school chemistry laboratory experiments, the salt bridge is usually a glass tube filled with an ionic solution with a porous material (cotton, glass wool) plugging the ends. The porous material allows ions to slowly move in or out of the tube. Another method used in labs is to have one of the electrodes inside a smaller, ceramic cup (**porous cup**) of solution that is placed right in the container of the other half cell. In household batteries, wet paper is used to separate the electrodes and a moist, pasty material acts as the solution to allow for the movement of ions.

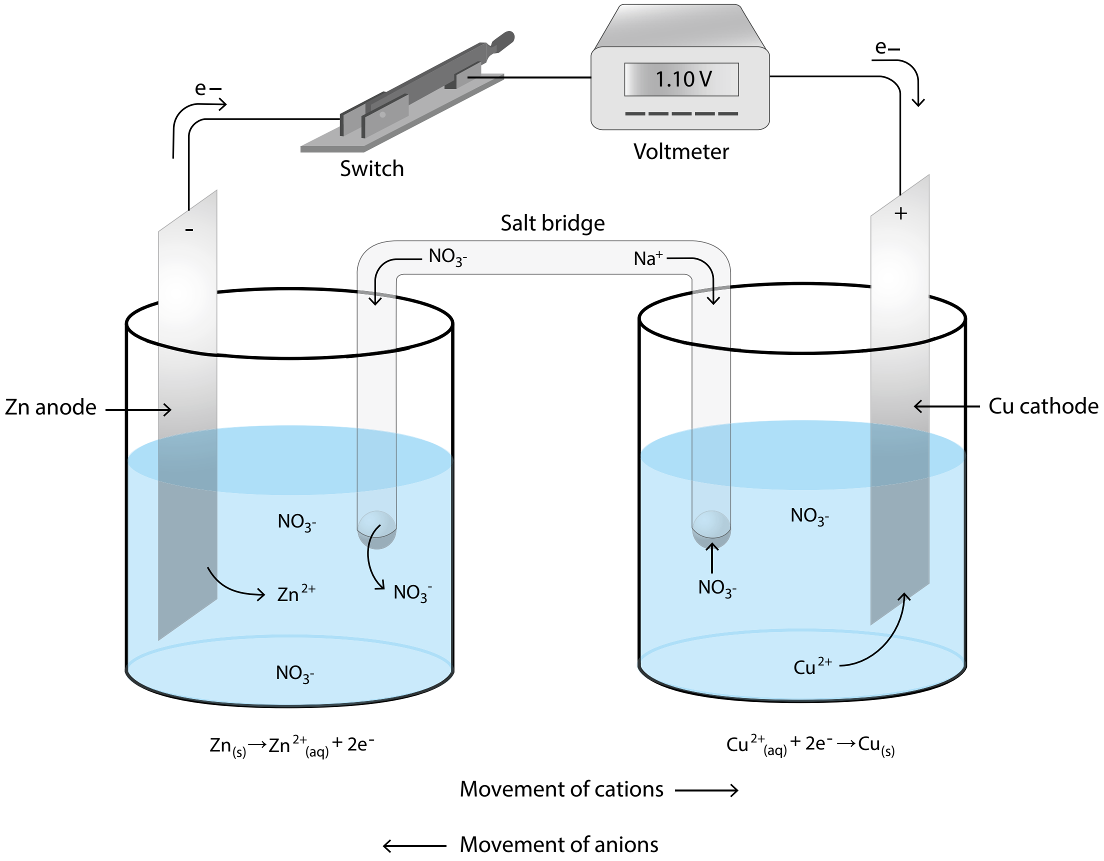
How an Electrochemical Cell Works

At this point, if you had access to the required materials, you would construct the cell and measure the voltage across the electrodes. In the diagram that follows, the two **half-cell reactions** are connected by a salt bridge. The following statements are true for all electrochemical cells.

- The electrode at which reduction occurs is termed the cathode. The charge on the cathode is positive.
- The electrode at which the oxidation occurs is called the anode. The charge on the anode is negative because it produces electrons.

- The electrons produced at the anode travel along the external wire to the cathode, where they are used to reduce the strongest oxidizing agent. This happens because the strongest oxidizing agent has a greater attraction for electrons and literally pulls them off the strongest reducing agent.
- The reaction in an electrochemical cell is always spontaneous.
- In the operating electrochemical cell, positive ions (cations) in the conducting solutions migrate toward the collection of electrons on the positive electrode, or cathode, while negative ions migrate (move) toward the negative electrode, or anode.

The following diagram illustrates the electrochemical cell you were introduced to in the previous lesson.



The following points summarize how this specific example of an electrochemical cell works.

1. Electrons are produced at the zinc strip according to the following oxidation reaction: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$.
2. Since zinc metal is oxidized, it is the anode (the negative electrode).

3. The electrons leave the zinc anode and pass through the external circuit to reach the copper strip. This flow of electrons can be controlled by a switch and measured by a voltmeter.

4. The copper strip accepts electrons, which interact with copper ions in solution according to this reduction reaction: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$.

5. Since copper ions are reduced at the copper strip, the copper strip is the cathode (the positive electrode).

6. To complete the circuit, both positive and negative ions move through the aqueous solution via the salt bridge. A salt bridge, which joins the two half cells, is filled with an electrolytic solution, such as potassium nitrate. The salt bridge allows for the movement of ions, without allowing the solutions to mix. The ions in the salt bridge do not need to match the ions in either half cell.

When the zinc nitrate and copper (II) nitrate solutions in the voltaic half-cell cells are both 1.0 mol/L concentrations, the cell generates an electrical potential of 1.10 volts.

Review the following animations to see more explanations pertaining to electrochemical cells.

Begin by viewing the following six animations explaining how a voltaic cell works:

- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt1.htm
- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt2.htm
- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt3.htm
- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt4.htm
- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt5.htm
- www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/Chapter17-Animations/ElectrochemCell_pt6.htm

You can also find these six animations on the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chem6.htm#electro.

A very clear explanation of various electrochemical cells is available online at www.funsci.com/fun3_en/electro/electro.htm.

A mouse rollover diagram showing the various parts of an electrochemical cell is available online at http://chimge.unil.ch/En/redox/1red13.htm.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Writing Half-Cell Reactions

There are a series of suggested steps to follow.

1. Write out the spontaneous reaction.
2. Identify the reduction reaction (one of the half-cell reactions).
3. Identify the oxidation reaction (the other half-cell reaction).
4. Confirm the spontaneity of the reaction.

At this point, if you had access to the required materials, you would then construct the cell and measure the voltage across the electrodes. If you were to do this, you would not likely achieve the predicted net E_{cell} voltage. The maximum voltage is dependent on concentration and assumes that temperature is constant. Immediately after the connections are made, concentrations will change. The reactant ions will decrease as they are used up and the product ions will increase as they are produced. Le Châtelier's principle will then cause a stress on the system and attempt to re-establish equilibrium by the reverse reaction and by reducing the net cell voltage. As the concentration of the reaction decreases, so does the voltage.

A simulation showing how to create an electrochemical cell is available online at http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/electroChem/voltaicCell20.html. You can select the electrodes and the ionic solutions needed to create a functioning electrochemical cell by opening and closing the tabs labelled Metals and Solutions.

If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

Learning Activity 6.3: The Lemon Battery Laboratory Activity

An example of an electrochemical cell is the simple lemon battery. You may have constructed such an apparatus in a science class or for a science fair. A small copper plate (or penny) and a small strip of zinc (a galvanized nail or a paper clip) are inserted into the flesh of a lemon. Using alligator clips and wire, these metal strips are connected to a small device like a digital clock or a small light bulb. The zinc metal undergoes oxidation, in this example. Apply your knowledge of electrochemical cells to answer the following questions about the lemon battery. If you have the means to construct this lemon battery, it would be beneficial to do so. If you do not have the required materials, you may use the following link to observe a similar reaction where lemon juice with zinc and copper electrodes are used to power a clock: www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/clock.htm.

You can also access this animation by

- going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chemvid8.htm#electro
- scrolling down the page until you find the video titled "***Electrochemical Cell Clock***"

1. Where are chemical reactions occurring in this electrochemical cell?
2. Identify the electrolyte in this electrochemical cell.
3. How do you know that the transfer of electrons does not proceed at the same rate in **both** directions?
4. Why would this light bulb **not** remain illuminated indefinitely?

Check the answer key.

Cell Potential (E^o)

The **cell potential** for a voltaic cell is a measure of the amount of energy that can be generated to drive an electric current through a wire. By definition, one **joule** of energy is produced when one **coulomb** of electrical charge is transported across a potential of one volt. Coulombs are the unit of measure of electric charge. One electron has a charge of -1.60×10^{-19} coulombs and one proton has a charge of $+ 1.60 \times 10^{-19}$ coulombs. One coulomb is equal to the charge on 6.24×10^{18} electrons. Electric charge can only flow between two electrodes when there is a difference in electric potential between the two points. This difference in potential is an indication of how much energy is available to move electrons from the anode to the cathode.

To better understand this concept, imagine a ball rolling down a hill. The ball will roll downhill to a low spot due to the difference in gravitational potential energy between the lower spot and the higher spot. As the ball rolls, it attains kinetic energy relative to the difference in height between the high spot and the low spot on the hill. This idea is similar to the energy of electrons flowing from an anode to a cathode in a voltaic cell. The voltage of the voltaic cell is determined by the difference in electric potential between the two electrodes, as one accepts electrons more readily than the other.

Cell potentials measured under standard-state conditions are represented by the symbol E^o. The greater the difference between the oxidizing and reducing strengths of the reactants and products, the greater the cell potential. To obtain a relatively large cell potential, a strong reducing agent must react with a strong oxidizing agent. For the purposes of this course, the following set of **standard-state conditions** for electrochemical measurements will be assumed.

You will notice that values for $E°$ are both negative and positive. A negative value for $E°$ means that there would be a net loss of energy if the half reaction proceeds. Consequently, the reduction reaction is non-spontaneous. A positive value for $E°$ means that the reduction is spontaneous. Also note that the reactions are reversible—which would represent the oxidation reaction. Reversing the reaction also reverses the value for $E°$ (inverting the sign).

- All solutions are 1 mol/L.
- All gases have a partial pressure of 1 atmosphere.
- Standard-state measurements are taken at 25^oC.

The following **Table of Standard Reduction Potentials with Values** lists the cell potentials for many different half-cells. Over time, chemists have measured and recorded the standard reduction potentials of many different half-cells. You will use this reference table to find the $E°$ values required for this module.

Table of Standard Reduction Potentials with Values

Oxidized Species	$\rightleftharpoons$	Reduced Species	E°(Volts)
$Li^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Li_{(s)}$	–3.04
$K^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$K_{(s)}$	–2.93
$Ca^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Ca_{(s)}$	–2.87
$Na^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Na_{(s)}$	–2.71
$Mg^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Mg_{(s)}$	–2.37
$Al^{3+}{}_{(aq)} + 3e^{-}$	$\rightleftharpoons$	$Al_{(s)}$	–1.66
$Mn^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Mn_{(s)}$	–1.19
$H_2O_{(l)} + e^{-}$	$\rightleftharpoons$	$\frac{1}{2}H_{2(g)} + OH^{-}{}_{(aq)}$	–0.83
$Zn^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Zn_{(s)}$	–0.76
$Fe^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Fe_{(s)}$	–0.45
$Ni^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Ni_{(s)}$	–0.26
$Sn^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Sn_{(s)}$	–0.14
$Pb^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Pb_{(s)}$	–0.13
$H^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Cu^{+}{}_{(aq)}$	+0.15
$SO_4{}^{2-}{}_{(aq)} + 4\ H^{+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^{-}$	$\rightleftharpoons$	$2OH^{-}{}_{(aq)}$	+0.40
$Cu^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^{-}$	$\rightleftharpoons$	$I^{-}{}_{(aq)}$	+0.54
$Fe^{3+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Fe^{2+}{}_{(aq)}$	+0.77
$Ag^{+}{}_{(aq)} + e^{-}$	$\rightleftharpoons$	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^{-}$	$\rightleftharpoons$	$Br^{-}{}_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^{+}{}_{(aq)} + 2e^{-}$	$\rightleftharpoons$	$H_2O_{(l)}$	+1.23
$Cr_2O_7{}^{2-}{}_{(aq)} + 14H^{+}{}_{(aq)} + 6e^{-}$	$\rightleftharpoons$	$2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^{-}$	$\rightleftharpoons$	$Cl^{-}{}_{(aq)}$	+1.36
$MnO_4{}^{-}{}_{(aq)} + 8H^{+}{}_{(aq)} + 5e^{-}$	$\rightleftharpoons$	$Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^{-}$	$\rightleftharpoons$	$F^{-}{}_{(aq)}$	+2.87

Writing Overall Reactions

We can use the Table of Standard Reduction Potentials with Values to determine the voltage of an electrochemical cell. At this point, you should be familiar with writing oxidation and reduction half-reactions from Module 1, using an activity series to predict reactions. You should also know that an electrochemical cell is made up of two half-cells, with oxidation occurring in one cell and reduction in the other cell.

Writing the overall reaction for an electrochemical cell involves writing out the net redox reaction and determining the cell potential (voltage) that can be produced by the reaction. You will learn more about how the half-cell potentials are determined in Lesson 4.

Use the following steps to write the overall reaction for an electrochemical cell:

Step 1: Determine the oxidation and reduction reactions and write out the half-reactions.

Step 2: Write the value for E° next to the reaction.

Because oxidation is the reverse of reduction, you will need to write the inverse of the value for E° when you reverse any reaction from the Table of Standard Reduction Potentials with Values. In other words, the oxidation reaction will have the opposite sign for E° from the table.

Step 3: Use coefficients to balance the electrons between the half reactions.

Step 4: Add the half reactions, cancelling electrons and spectator ions (if present) and add the values for E° to obtain the overall cell potential.

Example 1

What would be the overall reaction and cell potential for an electrochemical cell made with magnesium and silver electrodes? The electrodes will be placed in 1 mole solutions containing Mg^{2+} or Ag^{+}, respectively.

Step 1: Using the "up and to the right" rule, we see that Ag+(aq) will oxidize Mg(s). Write out the silver reaction as a reduction and the magnesium as an oxidation.

(Reduction) $Ag^{+}{}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$

(Oxidation) $Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^{-}$

Step 2: Write the value for E° next to the reaction (inverting the sign for the oxidation half-reaction).

(Reduction) $Ag^{+}{}_{(aq)} + e^{-} \rightarrow Ag_{(s)} \quad E°_{Red} = + 0.80 \text{ V}$

(Oxidation) $Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^{-} \quad E°_{Ox} = + 2.37 \text{ V}$

Step 3: Use coefficients to balance the electrons in the half reactions.

In this case, we multiply the reduction reaction through by 2.

$$2Ag^{+}{}_{(aq)} + 2e^{-} \rightarrow 2Ag_{(s)} \qquad E^{\circ}_{Red} = +\ 0.80\ V$$

$$Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^{-} \qquad E^{\circ}_{Ox} = +\ 2.37\ V$$

> **Note**
>
> Changing the coefficients on these reactions does not change the value for E°. E° is only dependent on the concentration of the reactants.

Step 4: Add the half reactions, cancelling electrons and spectator ions (if present) and adding the values for E° to obtain the overall cell potential.

$$2Ag^{+}{}_{(aq)} + \cancel{2e^{-}} \rightarrow 2Ag_{(s)} \qquad E^{\circ}_{Red} = +\ 0.80\ V$$

$$Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + \cancel{2e^{-}} \qquad E^{\circ}_{Ox} = +\ 2.37\ V$$

$$2Ag^{+}{}_{(aq)} + Mg_{(s)} \rightarrow 2Ag_{(s)} + Mg^{2+}{}_{(aq)} \qquad E^{\circ}_{Cell} = +\ 3.17\ V$$

Learning Activity 6.4: Writing Overall Reactions

1. In the previous example, which half-cell is the anode? How do you know?
2. In the previous example, which half-cell is the cathode? How do you know?

Check the answer key.

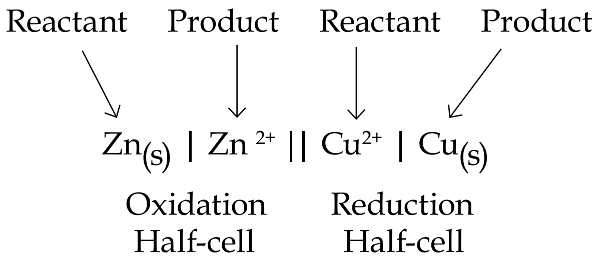
Short-hand Notation

A short-hand notation is often used to represent the voltaic cell diagram. There are three simple steps to writing this type of notation.

1. Write the reactant and product of the **oxidation** half-reaction on the **left-hand side**. A single line separates the two species.
2. Insert two vertical lines to represent the **salt bridge**.

3. Write the reactant and product of the **reduction** half-reaction on the **right-hand side**. A single line separates the two species.

Following is an example of short-hand notation, using the reaction from the previous section.



- The single vertical line represents a phase boundary between the metal and the ion in solution.
- The double vertical line represents the salt bridge.
- The anode is written first, to the left of the double lines.
- The cathode reaction is written second, to the right of the double lines.

Learning Activity 6.5: Understanding Voltaic Cells

1. For each example, draw a labelled electrochemical cell indicating the
   - electrodes
   - types of ions and their direction of motion
   - electrode reactions

   a. $3\ Mg_{(s)} + 2\ Al^{3+}_{(aq)} \rightarrow 3\ Mg^{2+}_{(aq)} + 2\ Al_{(s)}$

   b. $Cu^{2+}_{(aq)} + Mn_{(s)} \rightarrow Cu_{(s)} + Mn^{2+}_{(aq)}$

   c. $Br_{2(g)} + 2\ I^{-}_{(aq)} \rightarrow 2\ Br^{-}_{(aq)} + I_{2(s)}$

2. For each of the following electrochemical cells, indicate the
   - identity of the cathode
   - identity of the anode
   - electrode reactions
   - net reaction

continued

Learning Activity 6.5: Understanding Voltaic Cells (continued)

a. nickel and silver electrodes
b. lead and zinc electrodes
c. magnesium and chlorine electrodes
d. sodium and manganese electrodes

3. For each of the following pairs, determine which species are the reactants.

a. $Zn \mid Zn^{2+} \parallel Pb^{2+} \mid Pb$ ________________________
b. $Pb \mid Pb^{2+} \parallel Cu^{2+} \mid Cu$ ________________________
c. $Cu \mid Cu^{2+} \parallel Ag^{+} \mid Ag$ ________________________

Check the answer key.

Lesson Summary

In this lesson, you learned how the voltaic cell operates. In addition, you learned that electrons flow from the anode to the cathode of an electrochemical cell. Finally, you were introduced to cell potential as a measure of the energy that a cell can produce to push a current through a wire by a chemical reaction. In the next lesson, you will calculate standard electrode potentials for an electrochemical cell, given the electrode potentials for the half-reactions of that cell.

Assignment 6.3: The Functioning Voltaic Cell (11 marks)

1. Answer the following questions about a silver-lead cell set-up. Use the sample voltaic cell provided in this lesson to help you draw and understand the operation of the silver-lead cell. Show the work to support your answers, including writing half-reactions where necessary.

   a. Identify which metal will undergo reduction and identify the cathode. *(2 marks)*

   b. Write the half-cell reactions and the net (overall) reaction. *(3 marks)*

   c. Which metal is the anode? *(1 mark)*

continued

d. In which direction are the electrons moving? *(1 mark)*

2. What are the main components of a voltaic cell? Describe each of their roles in the operation of the cell. *(4 marks)*

LESSON 4: STANDARD ELECTRODE POTENTIAL (2.5 HOURS)

Lesson Focus

SLO C12-6-06: Define standard electrode potential. Include: hydrogen electrode as a reference.

SLO C12-6-07: Calculate standard cell potentials, given standard electrode potentials.

SLO C12-6-08: Predict the spontaneity of reactions using standard electrode potentials.

Lesson Introduction

In the previous lesson, you were introduced to the concept of standard reduction potential. In this lesson, you will continue to work with E°, including calculating the standard electrode potential of an electrochemical cell and predicting the spontaneity of reactions using standard electrode potentials for half-cells.

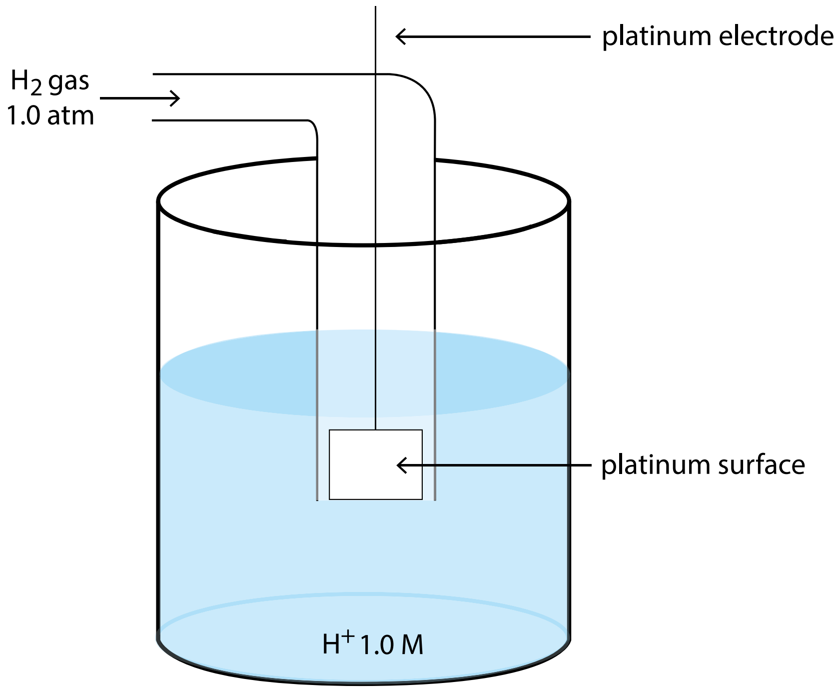
Electrode Potential

When scientists first constructed voltaic (galvanic) electrochemical cells, they recorded **standard reduction potentials** (E°) that resulted from the reactions; however, they did not know how much each half-cell contributed to the total net cell voltage. Many experiments were done in an attempt to determine the absolute E° for any half-cell reaction. Previously, you were introduced to a term that is often used synonymously with cell potential, electromotive force, abbreviated as **EMF**.

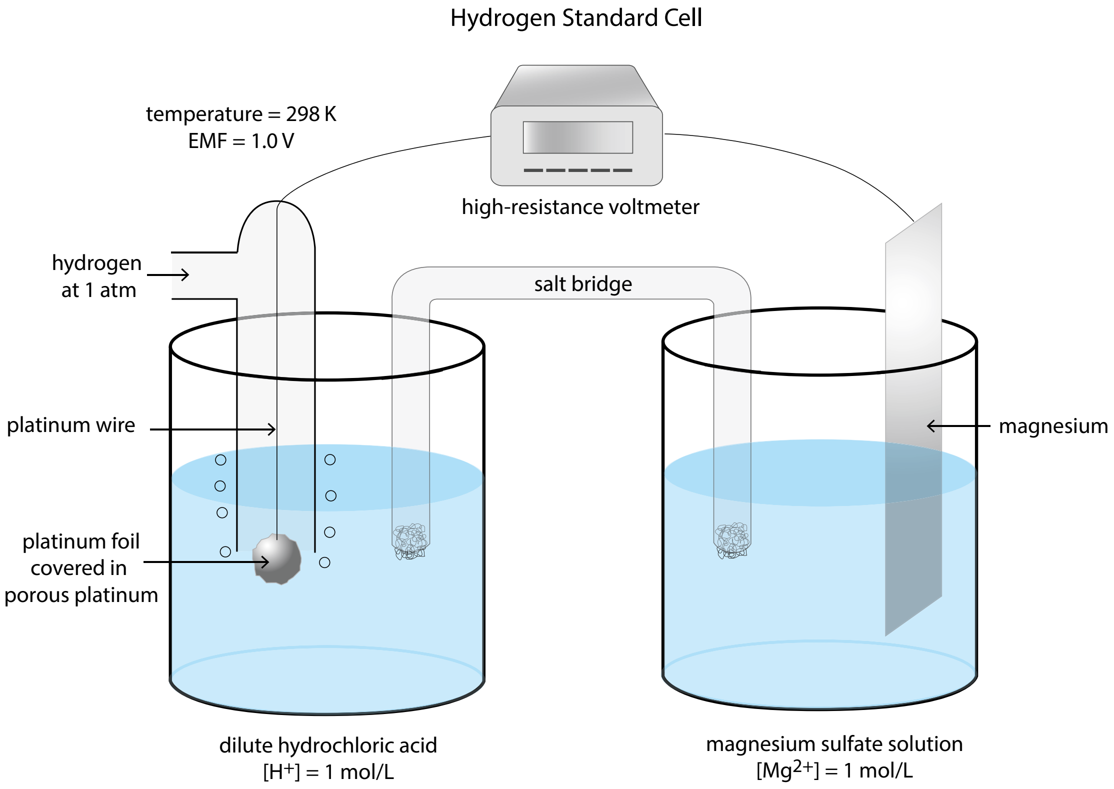
The values in the table are based on using the half-cell reaction that is being measured as the cathode and a standard hydrogen electrode as the anode. Many "reference" electrodes were tried before the hydrogen half-cell was chosen as the standard to which all other electrodes would be measured. When you look at the Table of Standard Reduction Potentials with Values, you will notice that the hydrogen half-cell reaction appears in the middle of the table and has an EMF value of 0.

In the hydrogen electrode shown below, hydrogen gas is bubbled into a hydrochloric acid solution at 25°C. While completing experiments, chemists found that not only did temperature affect the net cell potential, but so too did the concentration of ions in solution and, if a gas was used, the pressure. The platinum electrode provides a surface on which the dissociation of hydrogen molecules can occur and serves as an electrical conductor to the external circuit.

Standard Hydrogen Electrode



The following diagram demonstrates how a hydrogen cell can be used as a standard by which another metal's electrode potential can be measured.



Under standard conditions of 1 atmosphere (atm) for hydrogen gas and 1 mol/L HCl, the potential for the reduction of H^+ at 25°C is taken to be exactly zero.

$$2H^+ (1 \text{ mol/L}) + 2e^- \rightarrow H_2 (1 \text{ atm}) \qquad E° = 0 \text{ V}$$

Once the standard half-cell had been chosen, scientists were able to use this cell to determine the E° values for many other half-cell reactions. Note that this choice was arbitrary and could have been any half-reaction for the purpose of comparing against other half-reactions. These values were placed on a table of half-cell reactions containing standard reduction potentials.

Using the Table of Standard Reduction Potentials with Values

The table is organized according to a substance's tendency to gain electrons (in order of increasing reduction potential). This tendency of a substance to gain electrons is its reduction potential. For this reason, all of the half-reactions on the Table of Standard Reduction Potentials with Values are written as reductions; however, in a redox reaction one species must be oxidized. The half-reaction that appears higher on the table is oxidized, and the half-reaction that appears lower on the table is reduced. For the half-reaction that represents oxidation, you must reverse the equation (this makes sense since the opposite of a reduction is an oxidation). In doing so, you must also change the sign of the E° value.

> **Note**
>
> For every redox reaction, the half-reaction that is more **positive** will proceed as a **reduction** reaction and the half-reaction that is more **negative** will proceed as an **oxidation** reaction.

The Table of Standard Reduction Potentials with Values is used to determine the spontaneity and electrical potential of a given cell. Any **positive cell potential** value determined by finding the difference between the cathode and anode half-reaction potentials will result in a **spontaneous redox reaction**. Any **negative cell potential** value will indicate a **non-spontaneous redox reaction**.

Calculating Standard Electrode Potentials

There are several ways to determine overall cell potential. In this section, all three ways will be explored. You are free to use any of these methods to arrive at the correct answer.

Once you have determined the E° values for each half-reaction, you can calculate the voltage for an electrochemical cell (E°_{cell}) by adding the two values. The following are two methods for the calculation of $E^\circ_{\mathbf{cell}}$:

$$E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red}$$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$ **(You don't flip the signs.)**

To calculate the cell potential, the suggested steps are as follows:

1. Find the half-reactions from the reduction potentials table.
2. Substitute the half-cell potentials into the equation.
3. Use your answer to determine if the reaction is spontaneous or not.

Example 1

Calculate the cell potential for a silver-copper cell.

Step 1: Find the half-reactions for silver and copper from the reduction potentials table.

$$Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)} \quad E° = +0.34\ V$$

$$Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)} \quad E° = +0.80\ V$$

Since Ag^{+} ions are more easily reduced than Cu^{2+} ions, the $Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$ half-reaction is the reduction half-reaction and $Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)}$ needs to be reversed to become the oxidation half-reaction.

In other words, the Ag^{+} has a greater affinity for electrons than the Cu^{2+} ion and, as a result, the reaction with the lower E° will become the oxidation reaction. When reversing any half-reaction, the sign of the reduction potential is also reversed. Thus, the half-reactions would be

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-} \quad E° = -0.34\ V \quad \text{oxidation}$$

$$Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)} \quad E° = +0.80\ V \quad \text{reduction}$$

Step 2: Substitute the half-cell potentials into the equation.

$$E°_{cell} = E°_{ox} + E°_{red} = (-0.34\ V) + (+0.80\ V) = +0.46\ V$$

Step 3: Any positive cell potential value determined by finding the difference between the cathode and anode half-reaction potentials will result in a spontaneous redox reaction. This cell has a potential of + 0.46 V and confirms its spontaneity.

Example 2

Will a galvanic cell constructed according to the following overall reaction produce electricity? (All solutions are at 1 mol/L.)

$$Zn_{(s)} + Al^{3+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Al_{(s)}$$

Step 1: Find the half reactions:

$$\text{(Reduction)} \quad Al^{3+}_{(aq)} + 3e^{-} \rightarrow Al_{(s)} \quad E°_{Red} = -1.66\ V$$

$$\text{(Oxidation)} \quad Zn_{(s)} \rightarrow Zn^{2-}_{(aq)} + 2e^{-} \quad E°_{Ox} = +0.76\ V$$

Step 2: Substitute the half-cell potentials into the equation:

$$E°_{cell} = E°_{red} + E°_{ox} = (-1.66\text{ V}) + (+0.76\text{ V}) = -0.90\text{ volts}$$

Step 3: Since the value for $E°_{cell}$ is negative, this cell is not spontaneous and will not produce voltage.

Example 3:

A cell is made using zinc and gold metal as electrodes.

1. What is the cathode and what is the anode?
2. What is the net reaction?
3. What is the line notation for the cell?
4. What is the cell potential?

Answer:

1. Determine the oxidation and reduction reactions. The cathode will be the substance that is easiest to reduce. The substance easiest to reduce will have the reduction potential that is the most positive. The reduction reactions for each are

   $Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)} \quad E° = +1.50\text{ V}$

   $Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)} \quad E° = -0.76\text{ V}$

   Since gold ions are the most easily reduced, gold metal is the cathode and zinc metal is the anode.

2. Determine the net reaction by writing the oxidation and reduction reactions, then balance electrons lost and gained.

   Reduction: $Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

   Oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

   Balance electrons lost and gained.

   Reduction: $[Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}] \times 2$

   Oxidation: $[Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-] \times 3$

   Now you can write the net reaction.

   $\mathbf{2\ Au^{3+}_{(aq)} + 3\ Zn_{(s)} \rightarrow 2\ Au_{(s)} + 3\ Zn^{2+}_{(aq)}}$

3. Write the line notation. Write the anode (reactant) and the anode electrolyte (product) first.

   $Zn_{(s)} \mid Zn^{2+}$

Write the double lines to represent the salt bridge, and then the cathode electrolyte (reactant) and the cathode electrode (product).

$\mathbf{Zn \mid Zn^{2+} \parallel Au^{3+} \mid Au}$

4. Determine the cell potential. Even though you multiply the oxidation and reduction reactions, do **not** multiply the voltage.

$$\mathbf{E°_{cell} = E°_{ox} + E°_{red} = (+0.76\ V) + (+1.50\ V) = +2.26\ V}$$

Learning Activity 6.6: Using the Standard Reduction Potentials Table

1. Write the E° voltages for each half-cell reaction, the overall cell reaction, and the net cell voltage. Indicate if the reaction will be spontaneous as written or not. Explain your answer.
   a. $Ag^{+} + 1\ e^{-} \rightarrow Ag$
      $Cu \rightarrow Cu^{+} + 1\ e^{-}$
   b. $Au^{3+} + 3\ e^{-} \rightarrow Au$
      $Cr \rightarrow Cr^{3+} + 3\ e^{-}$
   c. $Cu^{+} + 1\ e^{-} \rightarrow Cu$
      $Al \rightarrow Al^{3+} + 3\ e^{-}$
   d. $Ca^{2+} + 2\ e^{-} \rightarrow Ca$
      $Fe \rightarrow Fe^{2+} + 2\ e^{-}$
   e. $2\ Br^{-} \rightarrow Br_2 + 2\ e^{-}$
      $Ni^{2+} + 2\ e^{-} \rightarrow Ni$
2. Complete the following reactions using the Table of Standard Reduction Potentials with Values. Determine the net cell potential and state if the reaction will occur.
   a. $Zn_{(s)} + Ag^{+}_{(aq)} \rightarrow$
   b. $Cu_{(s)} + Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} +$
   c. $Sn_{(s)} + Fe^{2+}_{(aq)} \rightarrow$
   d. $Mg_{(s)} + Sn^{2+}_{(aq)} \rightarrow$
3. What was the importance of finding the standard hydrogen electrode (SHE)?

Check the answer key.

Lesson Summary

In this lesson, you learned how to use the Table of Standard Reduction Potentials with Values. Using the table to determine the E° values for given half-reactions, you calculated the voltage for an electrochemical cell ($E°_{cell}$) by adding the two values. In the next lesson, you will learn about the operation of an electrolytic cell. This knowledge will allow you to compare and contrast voltaic (galvanic) and electrolytic cells.

Assignment 6.4: Calculating Standard Reduction Potentials (15 marks)

1. Complete the following reactions based on the Table of Standard Reduction Potentials. Calculate the net (overall) cell potential and indicate the reason the reaction proceeds spontaneously (or not) as written. You will receive 1 mark for completing the reaction, 1 mark for correctly calculating the E° value, 1 mark for stating the spontaneity, and 1 mark for explaining the spontaneity. *(4 marks per question x 3 questions = 12 marks total)*

   a. $Zn^{0}{}_{(s)} + Pb^{2+}{}_{(aq)} \rightarrow$

   b. $Cu^{0}{}_{(s)} + 2Ag^{+}{}_{(aq)} \rightarrow$

continued

Assignment 6.4: Calculating Standard Reduction Potentials (continued)

c. $Mn^0_{(s)} + 2\ Ca^{2+}_{(aq)} \rightarrow$

2. Explain how to use the Table of Standard Reduction Potentials with Values to determine the voltage of a voltaic cell. *(3 marks)*

Lesson 5: Comparing Voltaic and Electrolytic Cells (2 hours)

Lesson Focus

SLO C12-6-09: Compare and contrast voltaic (galvanic) and electrolytic cells.

SLO C12-6-10: Explain the operation of an electrolytic cell at the visual, particulate, and symbolic levels.
Include: a molten ionic compound and an aqueous ionic compound

SLO C12-6-11: Describe practical uses of electrolytic cells.
Examples: electrolysis of water, electrolysis of brine, electroplating, production and purification of metals...

Lesson Introduction

Up to this point, you have been working with electrochemical cells that convert electrical energy in a spontaneous reaction, thus forcing an electric current. In this lesson, you will be introduced to a second type of cell, known as the electrolytic cell. In doing so, you will study the conversion of electrical energy into chemical energy. Finally, you will learn about some practical uses of electrolytic cells.

Spontaneous versus Non-spontaneous

Electrochemical cells, such as the voltaic cell, are powered by redox reactions, which are spontaneous. These spontaneous redox reactions produce electrical energy and do not require that a current be supplied. The reverse reaction in each case is non-spontaneous and does require electrical energy in order for the reaction to occur. The general form of the reaction can be written as

Spontaneous →

Reactants ⇌ Products + Electrical Energy

← Non-spontaneous

To reverse the electrochemical cell, electrical energy must be provided for the redox reaction to proceed. Cells created in this way are **non-spontaneous** and are known as **electrolytic cells**. These cells are widely used to produce certain metals like sodium and aluminum from their oxides or ores, and also to electroplate gold and silver onto rings and other jewellery.

To explain what happens in an electrolytic cell, examine the decomposition of molten sodium chloride into sodium metal and chlorine gas. The reaction is as follows:

Non-spontaneous (electrolytic cell) →

$$2\ NaCl_{(l)} \leftrightarrow 2\ Na_{(s)} + Cl_{2(g)}$$

← Spontaneous (electrochemical cell)

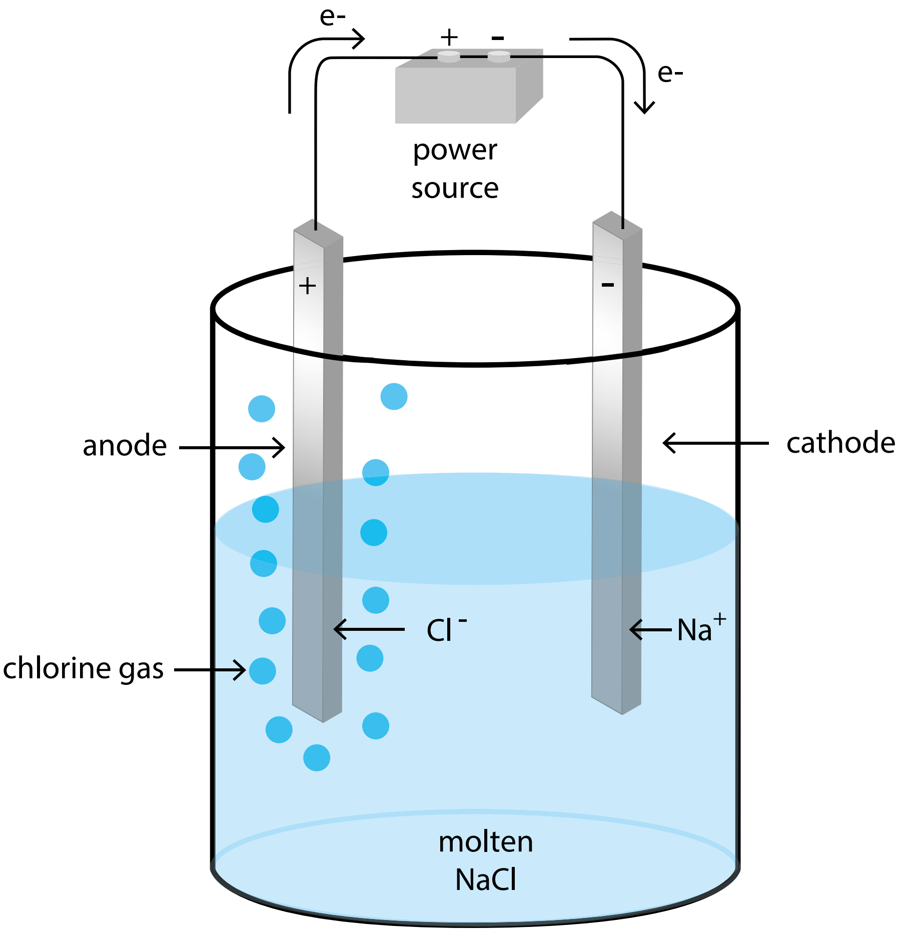
A labelled diagram of an electrolytic cell has been provided on the following page. Note the similarities to an electrochemical cell. One of the differences in the set-up of the apparatus is that in an electrolytic cell, both reactions occur in the same container and a source of electricity is required to push the usually non-spontaneous reaction to occur. The following items should be included in the diagram of an electrolytic cell: a container, an electrolytic solution (acid, base, or salt), two electrodes, an external electron “pump” (battery), a positive battery electrode connected to the anode, and a negative battery electrode connected to the cathode. Half-cell reactions are still necessary, as is the net reaction.

Molten Electrolytic Cells

The simplest electrolytic cells are those containing **molten ionic compounds**. When an ionic compound is melted, the ions separate from the crystal and flow freely. The dissociation is similar to the dissociation in water, except water is NOT present. For example, molten sodium chloride contains sodium ions and chloride ions in the liquid state, not to be confused with liquid sodium chloride.

$$NaCl_{(l)} \rightarrow Na^{+}_{(l)} + Cl^{-}_{(l)}$$

The electrolysis of molten sodium chloride decomposes $NaCl_{(s)}$ into liquid sodium and chlorine ions. This reaction is accomplished by heating sodium chloride beyond its melting point, 801°C, and by adding inert carbon electrodes, as shown in the simplified diagram that follows.



Using an external electrical source, the electrons are pushed onto the cathode. In the process of pushing electrons onto the cathode, electrons are drawn from the anode to replace those lost by the electrical source. The cations, Na^+ ions, move toward the cathode and become reduced. The anions, Cl^- ions, move toward the anode and become oxidized.

Note that the poles of the cell have changed. Since an external electrical source is used, the cathode is now the electron source and is, therefore, the negative electrode. The anode is the positive electrode. The reactions that occur are

Anode (oxidation)	$2\,Cl^-{}_{(l)} \rightarrow Cl_2{}^0{}_{(g)} + 2e^-$
Cathode (reduction)	$2\,Na^+{}_{(l)} + 2e^- \rightarrow 2\,Na^0{}_{(l)}$
Net reaction	$2\,Na^+{}_{(l)} + 2\,Cl^-{}_{(l)} \rightarrow 2\,Na^0{}_{(l)} + Cl_2{}^0{}_{(g)}$

The minimum voltage required for this cell to operate is given by the same equation as electrochemical cells.

$$E°_{cell} = E°_{ox} + E°_{red} = (-1.36V) + (-2.71V) = -4.07V$$

A power source with a minimum of 4.07 volts is needed to operate this cell. The negative potential indicates a non-spontaneous reaction. The industrial form of this cell is called a **Down's cell**. The Down's cell is the major method of producing sodium metal and pure supplies of chlorine gas.

Note that there are two different types of electrolytic cells. The more straightforward cell occurs when electricity is applied to a molten solution. The second type of cell occurs when electricity is applied to an **aqueous solution**. This cell is made more complicated by both the presence of water and the fact that there are many more species that can be oxidized and reduced.

Learning Activity 6.7: Comparing Voltaic and Electrolytic Cells

Complete the following table comparing characteristics of voltaic and electrolytic cells.

	Electrochemical Cell	**Electrolytic Cell**
Reaction Spontaneity		
Net Cell Potential		
Electricity Produced or Consumed		
Electrode Charge at the Anode and Cathode		
Location of Oxidation and Reduction		
Change in Energy		

Check the answer key.

The Process of Electrolysis

As the name suggests, **electrolysis** is the process of using electricity to break apart substances (electro = electricity, lysis = break). An electrolytic cell is a type of electrochemical cell that uses electricity to force a nonspontaneous reaction.

An electrolytic cell uses a direct current source attached to inert electrodes, usually carbon or platinum. These electrodes are inserted into the reactants. As with all electrochemical cells, oxidation occurs at the anode, and reduction at the cathode. The electrons still flow from the anode to the cathode; however, in an electrolytic cell they are "pushed" through the cell by an external electrical power source.

Many metals are refined using electrolysis. Demonstrations showing the refining of metals by electrolysis can be found by using a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

A simulation where you can create an electrolytic cell using different types of metals and various ionic solutions is available online at http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/electroChem/electrolysis10.html. Voltage and current can be adjusted as can the mass of the electrode.

> If you want to see other examples, go to a video hosting site such as www.youtube.com/ or a search engine such as www.google.ca/.

The Electrolysis of Water

The electrolysis of water is the decomposition of water into hydrogen and oxygen gas using an electric current. This is accomplished by placing inert electrodes into water and using an external power source to run an electric current through the water. Since water is a poor electrolyte, this process is very slow unless a small amount of electrolyte, such as H_2SO_4, is added.

In the decomposition reaction, water is reduced at the cathode and oxidized at the anode. You will find these in your Table of Standard Reduction Potentials. The half-cell reactions are

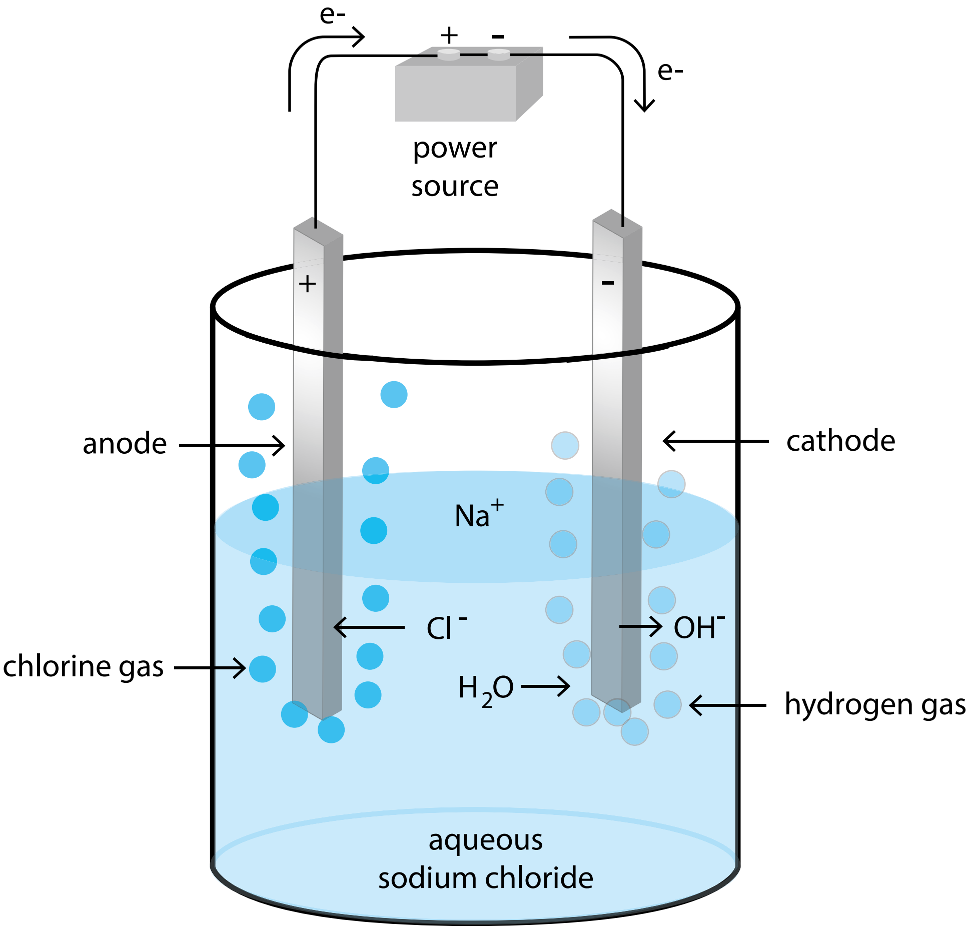
$$\begin{array}{ll} \text{Oxidation} & 2\,H_2O_{(l)} \rightarrow 4\,H^+_{(aq)} + O_{2(g)} + 4e^- \\ \text{Reduction} & 4\,H_2O_{(l)} + 4e^- \rightarrow 4\,OH^-_{(aq)} + 2\,H_{2(g)} \\ \hline \text{Net Reaction} & 6\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)} + 4\,H^+_{(aq)} + 4\,OH^-_{(aq)} \end{array}$$

Note that the 4 H^+ and the 4 OH^- in the products of the net reaction combine to produce 4 H_2O. This eliminates 4 waters from the reactants. The net equation is then $2\ H_2O_{(l)} \rightarrow 2\ H_{2(g)} + O_{2(g)}$.

Commercially, this is not a practical method for obtaining hydrogen gas or oxygen gas. Oxygen gas is more practically obtained from the air, and hydrogen can be produced from several sources such as hydrocarbons. A more commercially viable alternative for the production of hydrogen gas is the electrolysis of a concentrated sodium chloride solution, or brine.

The Electrolysis of Brine

Brine is a concentrated sodium chloride solution. The electrolysis of brine is a very valuable commercial process. The process occurs according to the following diagram.



The chloride ions are not easily reduced; therefore, at the cathode, water and sodium ions compete for electrons.

- Water is easier to reduce, so the products at the cathode are hydrogen gas and hydroxide ions.
- Sodium ions are not reduced to sodium metal in the process since water molecules are more easily reduced than sodium chloride. As a result, sodium is a spectator ion and does not need to be included in the overall reaction.
- At the anode, you would expect that water would be oxidized as it has a more positive oxidation potential. However, this is an exception to the rule: chlorine gas is actually produced at the anode. The reason for this unexpected result is called **overvoltage**. (The causes of overvoltage are quite complex and beyond the scope of this course.) Chlorine is produced at the anode rather than water because it is easier for the chloride ions to transfer their electrons to the electrode than it is for water, even though their oxidation potential does not suggest this.

Oxidation: $2\,Cl^-_{(aq)} \rightarrow Cl_{2(s)} + 2e^-$ (at anode)

Reduction: $2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$ (at cathode)

Overall reaction: $2\,Cl^-_{(aq)} + 2\,H_2O_{(l)} \rightarrow Cl_{2(s)} + H_{2(g)} + 2\,OH^-_{(aq)}$

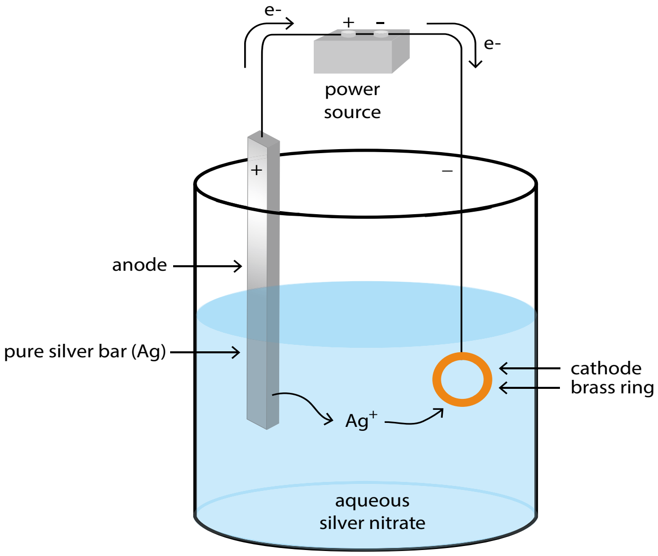
As the chloride ions are used up, sodium ions and hydroxide ions remain in solution. The electrolysis of brine is the major method for producing sodium hydroxide, chlorine gas, and very pure hydrogen gas. It is a more economical method of producing chlorine gas than molten sodium chloride since the mixture does not need to be heated to 801°C.

Lesson Summary

In this lesson, you learned of some practical uses of electrolytic cells, including electrolysis of water, electrolysis of brine, and molten electrolytic cells. In the next, you will use Faraday's Law to solve problems related to electrolytic cells.

Assignment 6.5: Electroplating (14 marks)

Conduct a short research on the electroplating process and its use. Using your research and the diagram showing the silver plating of a brass ring below, answer the following questions.



1. What is the main purpose of electroplating metals? *(2 marks)*

2. Give two examples of metals that might be electroplated. *(2 marks)*

continued

Assignment 6.5: Electroplating (continued)

3. Describe the general process of electroplating, including the positive and negative poles of the power source, the anode and the cathode, and the solution of metal ions. *(3 marks)*

4. Using the diagram on the previous page, describe the process of electroplating the brass ring. *(2 marks)*

5. What factors can be varied to control the thickness of the plating effect? *(1 mark)*

6. Write the reactions that occur at the anode and the cathode when plating a brass ring with silver. Identify which reaction represents the oxidation reaction and which reaction represents the reduction reaction. *(4 marks)*

LESSON 6: FARADAY'S LAW (3 HOURS)

Lesson Focus

SLO C12-6-12: Solve problems related to electrolytic cells, using Faraday's law.

Lesson Introduction

In the previous lesson, you were introduced to the electrolytic cell. In this lesson, the last of Module 6, you will learn about a mathematical law, called Faraday's law, that relates to these cells.

Faraday's Law

In every electrochemical process, whether spontaneous or not, a certain amount of electric charge is transferred during the oxidation and reduction. The half-reactions that represent electrode processes include the electrons that carry that charge. It is possible to measure the rate at which the charge is transferred with a device called an **ammeter**. An ammeter measures the current flowing through a circuit. The units of current are **amperes (A**). An ampere is defined as 1 coulomb flowing through a location on a conductor in 1 second. Remember that coulombs are the unit of measure for electric charge.

Unlike a voltmeter, ammeters allow electrons to pass and essentially "clock" them as they go by. The amount of electric charge that has passed through the circuit can then be calculated by these simple relationships.

Amperage = coulombs / second (or, rearranged as):

$$Q = I\Delta t$$

Q = charge (coulombs)
I = current (amperes)
Δt = change in time (seconds)

Charge = current x time **or** coulombs = amperes x seconds

These relationships connect reaction stoichiometry to electrical measurements. The principles underlying these relationships were worked out in the first half of the nineteenth century by English scientist **Michael Faraday**.

Michael Faraday (1791–1867), a British chemist, was the son of a blacksmith. He had no formal training in science. Instead of following his father's career, in 1804 he became an apprentice to a bookbinder. Many of the books he bound came from the local university. He read every book he bound, especially books on science.

Faraday developed a keen interest in science. He did not enjoy the bookbinding business but, with the help of friends and good luck, he was able to join the lab of Sir Humphry Davy. From there, he began to pursue his interest in electricity.

Faraday is the one who coined the terms *electrode, anode, cathode, electrolyte, ion, anion,* and *cation*.

In his work with electrolytic cells, in about 1833, Faraday discovered that **the amount of product formed at each electrode is directly proportional to the amount of electricity that passes through the cell.** This is called Faraday's law.

In recognition of Michael Faraday's pioneering work in this field of electrolysis, an important measurement was named after him. One **Faraday** is equal to the electric charge on one mole of electrons and, because electricity is made up of electrons, one Faraday represents one mole of electrons. (Another term used for a mole of electrons is a **gram equivalent**.) This important value is also called **Faraday's constant (F)** and is calculated as follows:

Faraday's constant (F)

$$= \text{Avogadro's number} \times \text{the charge on an electron}$$

$$= (6.02\text{x}10^{23}\cdot\text{mol}^{-1}) \text{ x } (1.602192\text{x}10^{-19}\text{coulombs}\cdot\text{electron}^{-1})$$

$$= 96{,}484 \text{ coulombs}\cdot\text{mole of electrons}^{-1}$$

$$\cong 96{,}500 \text{ C/mol}$$

With Faraday's constant, we can find the charge (Q) on any number of moles of electrons (n):

$$Q = nF$$

Combining the relationships between charge, amperage, and Faradays provides a way to calculate how much electricity is required to produce a mole of product at a given electrode.

Charge and current: $Q = I\Delta t$

Charge and Faradays: $Q = nF$

Combining these two equations: $nF = I\Delta t$

Rearranging we get

Faraday's law:

$$n = \frac{I\Delta t}{F}$$

Where,

n = moles of electrons

I = current (amperes)

Δt = time (seconds)

F = Faraday's constant (96500)

The units for Faraday's constant can be
96500 C/mol e^- or 96500 A·s/mol e^-

One Faraday will reduce one mole of sodium ions, since one electron is required to reduce one sodium ion.

$$Na^+_{(l)} + 1e^- \rightarrow Na_{(l)}$$

One Faraday will reduce one-half of a mole of magnesium ions, since one mole of magnesium is reduced by 2 electrons.

$$Mg^{2+}_{(l)} + 2e^- \rightarrow Mg_{(l)}$$

One Faraday will reduce one-third of a mole of aluminum ions, since one mole of magnesium is reduced by 2 electrons.

$$Al^{3+}_{(l)} + 3e^- \rightarrow Al_{(l)}$$

Alternatively, 3 Faradays are needed to reduce a mole of aluminum ions.

Notice that with Mg^{2+}, twice as many moles of electrons (electricity) are required to discharge 1 mole of Mg as opposed to 1 mol of Na. Since there is 1 mole of electrons in a Faraday, only half a mole, or 12.2 g, of Mg metal is deposited. It requires 1 Faraday to discharge 1 mole of Na^+ ions, 2 Faradays of electricity to discharge 1 mole of Mg^{2+} ions, and 3 Faradays to discharge 1 mole of Al^{3+} ions.

Solving Problems Using Faraday’s Law

There are many variables you can solve for using Faraday’s law. You can determine

- moles of electrons
- mass of product
- amount of time required for the reaction
- amount of current required to plate a given mass

Calculating Moles of Electrons

Example 1

If 7.85 amps flows through copper (I) chloride for 45.0 minutes, how many moles of electrons flow though the cell?

Half-reaction $Cu^{+} + 1e^{-} \rightarrow Cu^{0}$

Substituting into Faraday’s Law:

$$n = \frac{I \cdot \Delta t}{F} = \frac{(7.85\ A)(45\ min\ X\ 60 s/min)}{96\ 500\ A{\cdot}s/mole\ e^{-}}$$

$$= 0.220 \text{ moles of electrons}$$

Example 2

Determine the number of moles of electrons supplied by a battery with a current of 0.100 A for 50.0 min.

Substitute values into Faraday’s law. Remember to convert the minutes into seconds.

$$\text{moles } e^{-} = \frac{I \cdot \Delta t}{96\ 500\ C/mol} = \frac{(0.100\ A)(50.0\ min\ x\ 60 s/min)}{96\ 500\ C/mol}$$

$$= 0.00311 \text{ moles } e^{-}$$

Calculating Mass of Product

Example 3

Calculate the mass in grams of zinc deposited if 5.00 moles of electrons pass through a zinc sulphate solution when, at the cathode, zinc ions turn into zinc solid.

This is just simple stoichiometry and doesn't require Faraday's law.

Step 1: Write the reduction half-reaction.

$$Zn^{2+} + 2\,e^{-} \rightarrow Zn$$

Step 2: Use the mole ratio from the reaction.

2 moles of electrons → 1 mole of zinc metal

5 mol → x mol

$$\begin{aligned} x &= 2.50 \text{ mol X } 65.4 \text{ g/mol} \\ &= 164.5 \text{ g} \\ &= 164 \text{ g (rounded to 3 sig figs)} \end{aligned}$$

Example 4

Calculate the mass of aluminum produced by 7.50 A passing through molten aluminum oxide for 6 hours, 20 minutes, and 10 seconds. The reduction half-reaction is as follows:

$$Al^{3+}{}_{(l)} + 3e^{-} \rightarrow Al_{(l)}$$

Step 1: Calculate the moles of electrons supplied. Remember to convert time to seconds.

$$\text{moles } e^{-} = \frac{I \cdot \Delta t}{96\,500 \text{ C/mol}}$$

$$= \frac{(7.50 \text{ A})(6 \text{ hr x } 60 \text{ min/hr x } 60 \text{ s/min } + 20.0 \text{ min x } 60 \text{ s/min } + 10 \text{ s})}{96\,500 \text{ C/mol}}$$

$$= 1.77 \text{ moles } e^{-}$$

Step 2: Calculate moles of aluminum ions reduced. Three moles of electrons are needed to reduce 1 mole of aluminum ions.

$$\text{moles } Al^{3+} = 1.77 \text{ moles } e^{-}\left(\frac{1 \text{ mole } Al^{3+}}{3 \text{ moles } e^{-}}\right) = 0.590 \text{ moles } Al^{3+}$$

Step 3: Calculate the mass of aluminum.

$$\text{mass} = (0.590 \text{ moles})(27.0 \text{ g/mol}) = 15.9 \text{ g of aluminum is produced}$$

Example 5

If 9.00 amp flows for 10.0 min through a molten silver fluoride solution, what mass of silver metal would be deposited at the cathode?

Step 1: Determine the cathode reaction.

$$Ag^{+}{}_{(aq)} + 1e^{-} \rightarrow Ag^{0}{}_{(s)}$$

Step 2: Use the stoichiometry of the reaction to find the mass of Ag metal.

1 mol of electrons produces 1 mol Ag metal

Substitute into Faraday's law to find the moles of electrons produced

$$\text{moles } e^{-} = \frac{I \cdot \Delta t}{F} = \frac{(9.00\text{ A})(10.0\text{ min X } 60\text{s/min})}{96500\text{ A·s/mole } e^{-}}$$

$$= 0.0560 \text{ moles } e^{-}$$

0.0560 moles e^{-} will reduce 0.0560 moles of Ag

Mass of Ag = 0.0560 mol X 107.9 g/mol = 6.04 g Ag (rounded to 3 sig figs)

Calculating Amount of Time

Example 6

How long will it take to use up all the Cr^{3+} ions in 400.0 mL of a 0.120 mol/L solution using a current of 1.50 A?

Step 1: Find moles of Cr^{3+} in the solution.

$$n = CV = (0.120\text{ mol/L})(0.400\text{ L}) = 0.048 \text{ moles of } Cr^{3+}$$

Step 2: Use stoichiometry to find moles of electrons, then insert the data into Faraday's law.

$$Cr^{3+}{}_{(aq)} + 3e^{-} \rightarrow Cr_{(s)}$$

$$\text{moles } e^{-} = 0.048 \text{ moles } Cr^{3+}\left(\frac{3 \text{ moles } e^{-}}{1 \text{ mole } Cr^{3+}}\right) = 0.144 \text{ moles } e^{-}$$

$$\text{moles } e^{-} = \frac{I \cdot \Delta t}{F}$$

Rearranging for Δt and substituting,

$$\Delta t = \frac{\text{moles } e^{-} \cdot F}{l} = \frac{(0.144 \text{ moles } e^{-})(96500\text{ A·s/mol})}{1.50\text{ A}} = 9264\text{ s}$$

$$= 9260\text{ s (rounded to 3 sig figs)}$$

It takes 9260 s to remove all of the chromium (III) ions from the solution.

Calculating the Amount of Current Required to Plate a Given Mass

Example 7

A spoon was electroplated with silver. After 5.0 minutes of current, the spoon was 0.523 g heavier than before the plating. How much current must have passed through the spoon to plate this amount of silver?

Step 1: Find the moles of Ag in 0.523 g.

$$\text{moles Ag} = \frac{\text{(mass Ag)}}{\text{(molar mass of Ag)}} = \frac{\text{(0.523 g)}}{\text{(107.87 g/mole Ag)}} = 0.004848 \text{ moles}$$

Step 2: Use stoichiometry to find the moles of electrons used.

Cathode reaction:

$$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$$

Since 1 mole of electrons is required for each mole of silver ion reduced, we will need 0.004848 moles of electrons to plate 0.004848 moles of silver.

Step 3: Calculate the current using Faraday's law (remember to change time to seconds).

$$\text{moles } e^{-} = \frac{I \cdot \Delta t}{F}$$

Rearranging for current and substituting,

$$I = \frac{\text{moles } e^{-} \cdot F}{\Delta t} = \frac{(0.004848 \text{ moles } e^{-})(96500 \text{ A}\cdot\text{s/mol})}{(5.0 \text{ min X } 60 \text{ s/min})} = 1.56 \text{ A}$$

$$= 1.6 \text{ A (rounded to 2 sig figs)}$$

Learning Activity 6.8: Problem Solving with Faraday's Law

1. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten MgF_2.

   Cathode (reduction): $Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}$

2. How long would an aqueous gold (III) chloride cell need to operate to plate 2.5 g of gold on a bracelet with a current of 2.5 A?

   Cathode (reduction): $Au^{3+}{}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

3. If 10.0 g of sulphur is deposited on an electrode in an electrolytic cell from silver sulphide, calculate the mass of silver deposited on the other electrode.

   Cathode (reduction): $Ag^{+}{}_{(aq)} + 1e^- \rightarrow Ag_{(s)}$

4. What products would be expected at each electrode? Write the half-reactions and calculate the voltage required to electrolyze the following:
   a. aqueous copper (II) fluoride
   b. molten magnesium chloride
   c. aqueous iron (II) sulphate
   d. molten silver bromide
   e. aqueous aluminum chloride

5. How many seconds would be needed to generate 3.00 moles of electrons from 10.0 amp of current?

Check the answer key.

Lesson Summary

In this lesson, you learned about the work of Michael Faraday. You used Faraday's law to solve various problems relating to electrolytic cells. This is the last lesson in this module and in the course. Congratulations on your hard work!

Assignment 6.6: Using Faraday's Law (14 marks)

1. If 7.85 amp flows through a molten solution of copper I chloride for 45.0 minutes, how many moles of electrons flow through the cell? *(3 marks)*

2. Using 2.50 moles of electrons, what mass of copper metal would be produced from molten copper (II) sulphate?
   Reduction: $Cu^{2+}_{(l)} + 2e^{-} \rightarrow Cu_{(l)}$ *(3 marks)*

continued

Assignment 6.6: Using Faraday's Law (continued)

3. If 9.00 amp flows for 10.0 min through an aqueous silver nitrate solution, what mass of silver metal would be formed?
   Reduction: $Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$ *(4 marks)*

4. What time would be needed to deposit 42.50 g of zinc metal from 5.00 amp of current through a molten solution of zinc (II) bromide?
   Reduction: $Zn^{2+}_{(l)} + 2e^{-} \rightarrow Zn_{(l)}$ *(4 marks)*

Remember that you must also complete Assignment 1.9: Researching Redox Reactions at this point. You started working on this assignment in Module 1. Following is the marking rubric for this assignment.

Written Research Report Marking Rubric					
0 mark	**1 mark**	**2 marks**	**3 marks**	**4 marks**	**5 marks**
General Requirements					
The report does not fulfill any of the five requirements.	The report fulfills one of the five requirements.	The report fulfills two of the five requirements.	The report fulfills three of the five requirements.	The report fulfills four of the five requirements.	■ The report includes a title page. ■ The report is word-processed. ■ The report is double-spaced. ■ The report includes an alphabetized bibliography. ■ At least five sources are cited in the bibliography.
Content/Information					
The report does not fulfill any of the five requirements.	The report fulfills one of the five requirements.	The report fulfills two of the five requirements.	The report fulfills three of the five requirements.	The report fulfills four of the five requirements.	■ The information presented is clear, accurate, and concise. ■ All relevant information is completely presented and easily understood. ■ The report includes all redox reactions being utilized by the application. ■ The report identifies all substances being oxidized and reduced, as well as any oxidizing and reducing agents. ■ The report describes how the relevant redox reactions are used by man/nature.

continued

Written Research Report Marking Rubric (continued)

0 mark	1 mark	2 marks	3 marks	4 marks
Organization				
The report does not fulfill any of the four requirements.	The report fulfills one of the four requirements.	The report fulfills two of the four requirements.	The report fulfills three of the four requirements.	■ The clear and concise introduction includes a description of the topic of interest. ■ The body follows a logical pattern. ■ There are smooth transitions between paragraphs. ■ The conclusion includes a brief summary of the topic's connection to redox chemistry.
Mechanics (Grammar, Spelling, and Punctuation)				
Frequent errors demonstrate no sign of proofreading.	Five or six errors found.	Three or four errors found.	One or two errors found.	No errors found. Great job!
0 mark	**1 mark**	**2 marks**	**3 marks**	
Visual Aids (Pictures, Charts, and Diagrams)				
The visual aids do not fulfill any of the three requirements.	The visual aids fulfill one of the three requirements.	The visual aids fulfill two of the three requirements.	■ Time and effort were obviously put forth in the preparation of the visual aids. ■ The visual aids used in the report are neat and organized. ■ The visual aids help make the application more easily understood.	
Creativity				
The report shows no creativity.		The report shows some creativity.		The creativity in the report is "above and beyond."

MODULE 6 SUMMARY

You have now completed Module 6 and the course. It is time to submit your assignments and write your final examination. Read the following directions and double-check that you have submitted all the required work. Then, read the details regarding the final examination.

Submitting Your Assignments

It is now time for you to submit your assignments from Module 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 6 assignments and organize your material in the following order:

- ☐ Cover Sheet for Module 6 (found at the end of the course Introduction)
- ☐ Assignment 6.1: Predicting Spontaneity
- ☐ Assignment 6.2: History of the Voltaic Cell
- ☐ Assignment 6.3: The Functioning Voltaic Cell
- ☐ Assignment 6.4: Calculating Standard Reduction Potentials
- ☐ Assignment 6.5: Electroplating
- ☐ Assignment 6.6: Using Faraday's Law
- ☐ Assignment 1.9: Researching Redox Reactions (from Module 1)

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

Final Examination

Congratulations, you have finished the course and it is time to complete the final examination. It is worth 100 marks, or 30 percent of your final mark. In order to do well on the final examination, you should review all of your learning activities and assignments from 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, 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: pens, blank paper, a ruler, and a calculator. A maximum of **3 hours** is available to complete your final examination. When you have completed it, the proctor will then forward it to the Distance Learning Unit for assessment. Good luck!

Final Practice Examination and Answer Key

To help you succeed in your examination, a practice examination can be found in the learning management system (LMS). The final practice examination is very similar to the actual examination that you will be writing. The answer key is also included so that, when you have finished writing the practice examination, you can check your answers. This will give you the confidence that you need to do well on your examination. If you do not have access to the Internet, contact the Distance Learning Unit at 1-800-465-9915 to get a copy of the practice examination and the answer key.

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GRADE 12 CHEMISTRY (40S)

Module 6
Electrochemistry

Learning Activity Answer Keys

MODULE 6:
ELECTROCHEMISTRY

Learning Activity 6.1: Using the Activity Series

Use the activity series table to complete the following questions:

1. Indicate whether or not the following reactions will occur based on their relative position on the activity series table. If the reaction occurs, complete the equation.

   a. $Cu^+ + Pb \rightarrow$
   $2\,Cu^+ + Pb \rightarrow 2\,Cu + Pb^{2+}$ (Pb above Cu)

   b. $Sn + Cr^{3+} \rightarrow$ **no reaction**

   c. $NO_3^- + 4\,H^+ + Fe^{2+} \rightarrow$
   $NO_3^- + 4\,H^+ + 3\,Fe^{2+} \rightarrow NO + 2\,H_2O + 3\,Fe^{3+}$

   d. $Pb^{2+} + Ag_2S \rightarrow$ **no reaction**

2. Answer the following questions about these species: Au^{3+}, $Cr^0_{(s)}$, Sn^{2+}, and Br^-.

   a. Which is the most easily reduced or, in other words, the strongest oxidizing agent?

   Au^{3+}

   b. Which has the greatest affinity for electrons?

   Au^{3+}

   c. Which is the least easily oxidized?

   Br^-

   d. Which is the most easily oxidized?

   Cr^0

   e. Which will oxidize Sn^{2+} to Sn^{4+}?

   Au^{3+}

   f. Which will reduce $F_{2(g)}$ to $2F^-_{(aq)}$?

   Br^- and Cr^0

3. Consider the following equations:

   K + LiCl → ?

   Li + KCl → ?

   Why can't they both occur spontaneously?

   They can't occur spontaneously since one metal atom must be more reactive than the other. One must have a stronger pull on or "win the competition for" the chlorine atom.

4. How can you determine which atom will have a stronger pull on, or "win the competition for," another atom? Provide an example.

   Not all elements are equal in their ability to replace other elements. Metals, such as lithium, sodium, potassium, strontium, and calcium, are referred to as "very active" metals due to the fact that they give up electrons very easily. These metals can replace hydrogen from cold water to produce hydroxide and hydrogen gas: $2K + 2H_2O \rightarrow 2KOH + H_2$.

   Somewhat "less active" metals, such as aluminum, magnesium, chromium, and zinc, will only react with hot water and will produce oxide and hydrogen gas: $Mg + H_2O \rightarrow MgO + H_2$.

   Metals such as lead and iron do not react with water, but can react with HCl to form hydrogen gas: $Pb + 2HCl \rightarrow H_2 + PbCl_2$.

   Some "less active" metals will not displace hydrogen from HCl but will displace other metal ions in aqueous solutions.

Learning Activity 6.2: History of the Voltaic Cell

1. You have been introduced to a short history of the voltaic cell in Lesson 6.2. Construct a timeline to outline some of the major events in electrochemistry. Be sure to include Luigi Galvani's discovery of "animal electricity," Volta's first battery, Gaston Plante's lead battery, and, finally, Waldmar Jungner's Ni-Cad battery. More information on the contribution of the four individuals to the voltaic cell can be found by conducting research.

   Luigi Galvani (1780–1786): He used nerve impulses to explain his theory of animal electricity.

   Volta (1800): He assembled alternating discs of zinc and copper with pieces of cardboard soaked in salt water between the metals. The voltaic pile produced electrical current. Alessandro Volta's voltaic pile was the first "wet cell battery" that produced a reliable, steady current of electricity.

Gaston Plante's Lead Battery (1859): He created the first lead cell rechargeable battery. The basic design is still used in cars today.

Waldmar Jungner (1899): He created the first nickel cadmium rechargeable battery.

Learning Activity 6.3: The Lemon Battery Laboratory Activity

An example of an electrochemical cell is the simple lemon battery. You may have constructed such an apparatus in a science class or for a science fair. A small copper plate (or penny) and a small strip of zinc (or a paper clip) are inserted into the flesh of a lemon. Using alligator clips and wire, these metal strips are connected to a small device like a digital clock or a small light bulb. The zinc metal undergoes oxidation, in this example. Apply your knowledge of electrochemical cells to answer the following questions about the lemon battery. If you have the means to construct this lemon battery, it would be beneficial to do so. If you do not have the required materials, you may use the following link to observe a similar reaction where lemon juice with zinc and copper electrodes are used to power a clock: www.dlt.ncssm.edu/core/Chapter17-Electrochemistry/clock.htm.

You can also access this animation by

- going to the NCSSM T.I.G.E.R. site at www.dlt.ncssm.edu/tiger/chemvid8.htm#electro
- scrolling down the page until you find the video titled *"Electrochemical Cell Clock"*

1. Where are chemical reactions occurring in this electrochemical cell?

   There is a chemical reaction between the zinc and the lemon juice as well as between the copper strip and the lemon juice. These two chemical reactions push electrons through the wires.

2. Identify the electrolyte in this electrochemical cell.

   Lemon juice, an acid, serves as the electrolyte.

3. How do you know that the transfer of electrons does not proceed at the same rate in **both** directions?

   The electrons get pushed harder in one direction than the other because the two metals are different. If the metals were the same, the push would be equal and no electrons would flow.

4. Why would this light bulb **not** remain illuminated indefinitely?

   Over time, the zinc electrode will corrode in the acidic environment and increase the concentration of zinc ion solution. As this happens, the reaction slows down and the voltage decreases. Also, the lemon can dry up resulting in decreased voltage.

Learning Activity 6.4: Writing Overall Reactions

1. In the previous example, which half-cell is the anode? How do you know?

   Oxidation occurs at the anode, where cations are produced. The Zn/Zn^{2+} half-cell is, therefore, the anode.

2. In the previous example, which half-cell is the cathode? How do you know?

   Reduction occurs at the cathode, where cations are produced. The Cu^{2+}/Cu half-cell is the cathode.

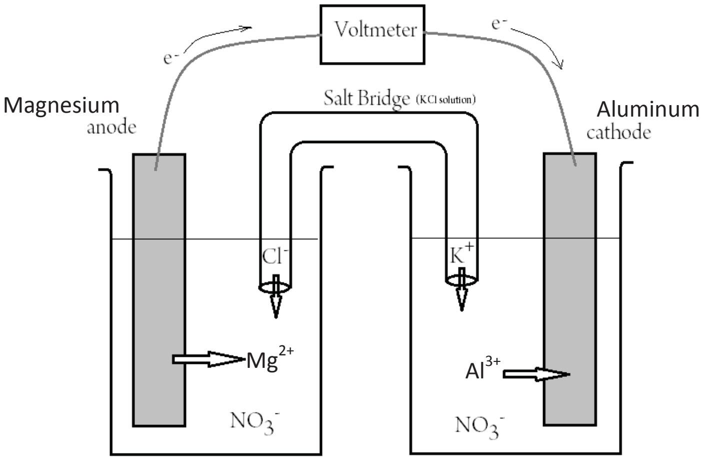
Learning Activity 6.5: Understanding Voltaic Cells

1. For each example, draw a labelled electrochemical cell indicating the
   - electrodes
   - types of ions and their direction of motion
   - electrode reactions

   a. $3\,Mg_{(s)} + 2\,Al^{3+}{}_{(aq)} \rightarrow 3\,Mg^{2+}{}_{(aq)} + 2\,Al_{(s)}$

   Anode (oxidation) $\mathbf{Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e-}$

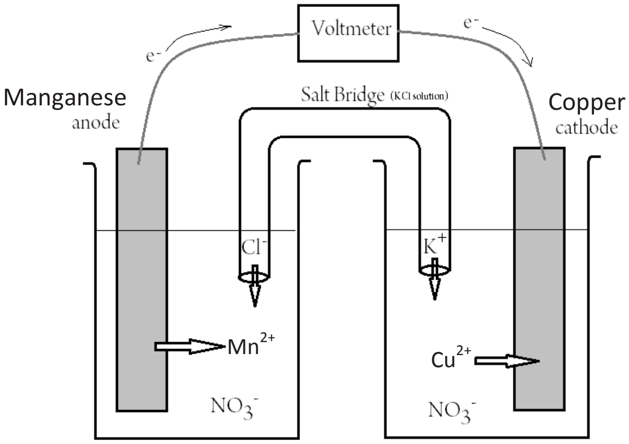
   Cathode (reduction) $\mathbf{Al^{3+} + 3e- \rightarrow Al_{(s)}}$



b. $Cu^{2+}_{(aq)} + Mn_{(s)} \rightarrow Cu_{(s)} + Mn^{2+}_{(aq)}$

Anode (oxidation) $Mn_{(s)} \rightarrow Mn^{2+}_{(aq)} + 2e^{-}$

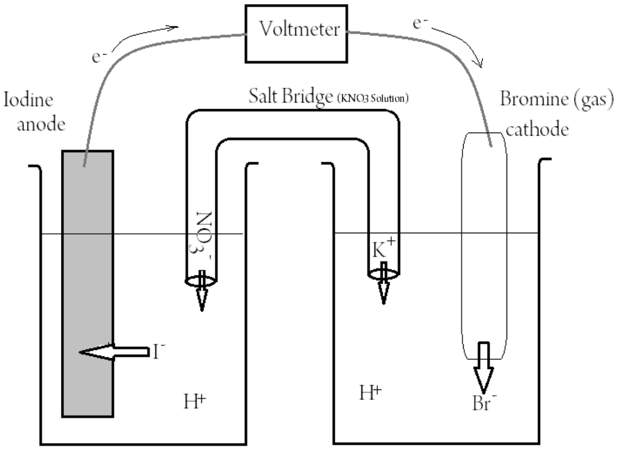
Cathode (reduction) $Cu^{2+} + 2e^{-} \rightarrow Cu_{(s)}$



c. $Br_{2(g)} + 2\,I^{-}_{(aq)} \rightarrow 2\,Br^{-}_{(aq)} + I_{2(s)}$

Anode (oxidation) $2\,I^{-}_{(aq)} \rightarrow I_{2(s)} + 2e^{-}$

Cathode (reduction) $Br_{2(g)} + 2e^{-} \rightarrow 2\,Br^{-}_{(aq)}$



2. For each of the following electrochemical cells, indicate the
   - identity of the cathode
   - identity of the anode
   - electrode reactions
   - net reaction

   a. nickel and silver electrodes

      $Ni^{2+}{}_{(aq)} + 2e^{-} \rightarrow Ni_{(s)}$ $E° = -0.25V$

      $Ag^{+}{}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$ $E° = +0.80V$

      — **Silver is more easily reduced (greatest tendency to attract electrons), so it is the cathode.**

      — **Nickel metal is the anode.**

      — **Anode (oxidation) $Ni_{(s)} \rightarrow Ni^{2+}{}_{(aq)} + 2e^{-}$**

      — **Cathode (reduction) $(Ag^{+}{}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}) \times 2$**

      — **$2\,Ag^{+}{}_{(aq)} + Ni_{(s)} \rightarrow 2\,Ag_{(s)} + Ni^{2+}{}_{(aq)}$**

b. lead and zinc electrodes

$Pb^{2+}{}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$ $E° = –0.13V$

$Zn^{2+}{}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ $E° = –0.72V$

- **Lead is more easily reduced (more positive *E*°), so it is the cathode.**
- **Nickel metal is the anode.**
- **Anode (oxidation) $Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e–$**
- **Cathode (reduction) $Pb^{2+}{}_{(aq)} + 2e– \rightarrow Pb_{(s)}$**
- **$Pb^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}{}_{(aq)}$**

c. magnesium and chlorine electrodes

$Mg^{2+}{}_{(aq)} + 2e^- \rightarrow Mg_{(s)}$ $E° = –2.37V$

$Cl_{2(g)} + 2e^- \rightarrow 2\,Cl^-{}_{(aq)}$ $E° = +1.36V$

- **Chlorine is more easily reduced (more positive *E*°), so it is the cathode.**
- **Magnesium metal is the anode.**
- **Anode (oxidation) $Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^-$**
- **Cathode (reduction) $Cl_{2(g)} + 2e^- \rightarrow 2\,Cl^-{}_{(aq)}$**
- **$Cl_{2(g)} + Mg_{(s)} \rightarrow 2\,Cl^-{}_{(aq)} + Mg^{2+}{}_{(aq)}$**

d. sodium and manganese electrodes

$Mn^{2+}{}_{(aq)} + 2e^- \rightarrow Mn_{(s)}$ $E° = –1.18V$

$Na^+{}_{(aq)} + 1e^- \rightarrow Na_{(s)}$ $E° = –2.71V$

- **Manganese is more easily reduced, so it is the cathode.**
- **Sodium metal is the anode.**
- **Anode (oxidation) $(Na_{(s)} \rightarrow Na^+{}_{(aq)} + 1e^-) \times 2$**
- **Cathode (reduction) $Mn^{2+}{}_{(aq)} + 2e^- \rightarrow Mn_{(s)}$**
- **$Mn^{2+}{}_{(aq)} + 2\,Na_{(s)} \rightarrow Mn_{(s)} + 2\,Na^+{}_{(aq)}$**

3. For each of the following pairs, determine which species are the reactants.

   a. $Zn \mid Zn^{2+} \parallel Pb^{2+} \mid Pb$ **Zn, Pb^{2+}**
   b. $Pb \mid Pb^{2+} \parallel Cu^{2+} \mid Cu$ **Pb, Cu^{2+}**
   c. $Cu \mid Cu^{2+} \parallel Ag^{+} \mid Ag$ **Cu, Ag^{+}**

Learning Activity 6.6: Using the Standard Reduction Potentials Table

1. Write the E° voltages for each half-cell reaction, the overall cell reaction, and the net cell voltage. Indicate if the reaction will be spontaneous as written or not. Explain your answer.

   a. $Ag^{+} + 1\,e^{-} \rightarrow Ag$

   $Cu \rightarrow Cu^{+} + 1\,e^{-}$

   $$\mathbf{Ag^{+} + 1\,e^{-} \rightarrow Ag \quad E^\circ = +0.80V}$$
   $$\mathbf{Cu \rightarrow Cu^{+} + 1\,e^{-} \quad E^\circ = -0.52V}$$

   $$\mathbf{Ag^{+} + Cu \rightarrow Ag + Cu^{+} \quad E^\circ = +0.28V}$$

   Spontaneous, cell potential is positive

   b. $Au^{3+} + 3\,e^{-} \rightarrow Au$

   $Cr \rightarrow Cr^{3+} + 3\,e^{-}$

   $$\mathbf{Au^{3+} + 3\,e- \rightarrow Au \quad E^\circ = +1.50V}$$
   $$\mathbf{Cr \rightarrow Cr^{3+} + 3\,e^{-} \quad E^\circ = +0.74V}$$

   $$\mathbf{Au^{3+} + Cr \rightarrow Au + Cr^{3+} \quad E^\circ = +2.24V}$$

   Spontaneous

   c. $Cu^{+} + 1\,e^{-} \rightarrow Cu$

   $Al \rightarrow Al^{3+} + 3\,e^{-}$

   $$\mathbf{3 \times (Cu^{+} + 1\,e^{-} \rightarrow Cu) \quad E^\circ = +0.52V}$$
   $$\mathbf{Al \rightarrow Al^{3+} + 3\,e^{-} \quad E^\circ = +1.66\ V}$$

   $$\mathbf{3\,Cu^{+} + Al \rightarrow 3\,Cu + Al^{3+} \quad E^\circ = +2.18V}$$

   Spontaneous

d. $Ca^{2+} + 2\,e^- \rightarrow Ca$

$Fe \rightarrow Fe^{2+} + 2\,e^-$

$Ca^{2+} + 2\,e^- \rightarrow Ca \quad E^\circ = -2.87V$

$Fe \rightarrow Fe^{2+} + 2\,e^- \quad E^\circ = +0.44V$

$Ca^{2+} + Fe \rightarrow Ca + Fe^{2+} \quad E^\circ = -2.43V$

Not spontaneous, cell potential is negative

e. $2\,Br^- \rightarrow Br_2 + 2\,e^-$

$Ni^{2+} + 2\,e^- \rightarrow Ni$

$2\,Br^- \rightarrow Br_2 + 2\,e^- \quad E^\circ = -1.06V$

$Ni^{2+} + 2\,e^- \rightarrow Ni \quad E^\circ = -0.25V$

$2\,Br^- + Ni^{2+} \rightarrow Br_2 + Ni \quad E^\circ = -1.31V$

Not Spontaneous

2. Complete the following reactions using the Table of Standard Reduction Potentials with Values. Determine the net cell potential and state if the reaction will occur.

a. $Zn_{(s)} + Ag^+_{(aq)} \rightarrow$

$(Ag^+_{(aq)} + 1\,e^- \rightarrow Ag_{(s)}) \times 2 \quad E^\circ = +0.80V$

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^- \quad E^\circ = +0.76V$

$2\,Ag+_{(aq)} + Zn_{(s)} \rightarrow 2\,Ag_{(s)} + Zn^{2+}_{(aq)} \quad E^\circ = +1.56V$

Spontaneous, cell potential is positive

b. $Cu_{(s)} + Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)}+$

$(Ag^+_{(aq)} + 1\,e^- \rightarrow Ag_{(s)}) \times 2 \quad E^\circ = +0.80V$

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^- \quad E^\circ = -0.34V$

$2\,Ag^+_{(aq)} + Cu_{(s)} \rightarrow 2\,Ag_{(s)} + Cu^{2+}_{(aq)} \quad E^\circ = +0.46V$

Spontaneous, cell potential is positive

c. $Sn_{(s)} + Fe^{2+}_{(aq)} \rightarrow$

$Fe^{2+}_{(aq)} + 2\,e^- \rightarrow Fe_{(s)} \quad E^\circ = -0.44V$

$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2\,e^- \quad E^\circ = +0.14V$

$Fe^{2+}_{(aq)} + Sn_{(s)} \rightarrow Fe_{(s)} + Sn^{2+}_{(aq)} \quad E^\circ = -0.30V$

Not spontaneous, cell potential is negative

d. $Mg_{(s)} + Sn^{2+}_{(aq)} \rightarrow$

$$\mathbf{Sn^{2+}_{(aq)} + 2\,e^{-} \rightarrow Sn_{(s)} \quad E^{\circ} = -0.14V}$$

$$\mathbf{Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2\,e^{-} \quad E^{\circ} = +2.37V}$$

$$\mathbf{Sn^{2+}_{(aq)} + Mg_{(s)} \rightarrow Sn_{(s)} + Mg^{2+}_{(aq)} \quad E^{\circ} = +2.23V}$$

Not spontaneous, cell potential is negative

3. What was the importance of finding the standard hydrogen electrode (SHE)?

   Any given pair of electrodes will give a specific cell E. To compare the relative strengths of voltaic half-cells, scientists needed to find a cell to which all others could be compared. Many reference cells were initially used, but the hydrogen cell conveniently fell in the middle of the table of potentials.

Learning Activity 6.7: Comparing Voltaic and Electrolytic Cells

Complete the following table comparing characteristics of voltaic and electrolytic cells.

	Electrochemical Cell	Electrolytic Cell
Reaction Spontaneity	**Spontaneous**	**Non-spontaneous**
Net Cell Potential	**Positive**	**Negative**
Electricity Produced or Consumed	**Produces**	**Consumes**
Electrode Charge at the Anode and Cathode	**Cathode +** **Anode –**	**Cathode –** **Anode +**
Location of Oxidation and Reduction	**Reduction – Cathode** **Oxidation – Anode**	**Reduction – Cathode** **Oxidation – Anode**
Change in Energy	**Converts chemical energy to electrical energy**	**Converts electrical energy to chemical energy**

Learning Activity 6.8: Problem Solving with Faraday's Law

1. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten MgF_2.

   Cathode (reduction): $Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}$

   $$\mathbf{n_{e^-} = \frac{I \cdot t}{96500\ C/mol} = \frac{(15.0\ A)(7.51\ min \times 60\ s/min)}{96500\ C/mol} = 0.07004\ moles\ e^-}$$

   Cathode (reduction): $\mathbf{Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}}$

   $$\mathbf{moles\ Mg = 0.007004\ moles\ e^- \left(\frac{1\ mole\ Mg}{2\ moles\ e^-}\right) = 0.0350\ moles\ Mg}$$

   $$\mathbf{m = n \bullet M = (0.0350\ moles)(24.3\ g/mol) = 0.851\ g\ Mg}$$

   Anode (oxidation): $\mathbf{2\ F^-{}_{(l)} \rightarrow F_{2(g)} + 2e^-}$

   $$\mathbf{moles\ F_2 = 0.007004\ moles\ e^- \left(\frac{1\ mole\ F_2}{2\ moles\ e^-}\right) = 0.0350\ moles\ F_2}$$

   $$\mathbf{m = n \cdot M = (0.0350\ moles)(38.0\ g/mol) = 1.33\ g\ F_2}$$

   0.851 g of Mg is formed at the cathode and 1.33 g of F_2 gas is formed at the anode.

2. How long would an aqueous gold (III) chloride cell need to operate to plate 2.5 g of gold on a bracelet with a current of 2.5 A?

   Cathode (reduction): $Au^{3+}{}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

   $$\mathbf{moles\ Au = \frac{m}{M} = \frac{2.50\ g}{197.0\ g/mol} = 0.01269\ moles\ Au}$$

   $$\mathbf{moles\ e^- = 0.01269\ moles\ Au \left(\frac{3\ moles\ e^-}{1\ mole\ Au}\right) = 0.03807\ moles\ e^-}$$

   $$\mathbf{t = \frac{(96500\ C/mol)n_{e^-}}{I} = \frac{(96500\ C/mol)(0.03807\ moles\ e^-)}{2.5\ A} = 1469.5\ s}$$

   It would require about 1470 s to plate 2.5 g of gold.

3. If 10.0 g of sulphur is deposited on an electrode in an electrolytic cell from silver sulphide, calculate the mass of silver deposited on the other electrode.

   Cathode (reduction): $Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$

   Find the moles of sulfur deposited.

   $$\textbf{moles S} = \frac{\textbf{m}}{\textbf{molar mass}} = \frac{\textbf{10.0 g}}{\textbf{32.1 g/mol}} = \textbf{0.3115 moles S deposited}$$

   Find moles of e⁻ needed.

   $$S^{2-} \rightarrow S + 2e^{-}$$

   $$\textbf{moles } e^{-} = \textbf{0.3115 moles S}\left(\frac{\textbf{2 moles } e^{-}}{\textbf{1 mole S}}\right) = \textbf{0.623 moles } e^{-}$$

   Find the moles of silver deposited from 0.623 moles of e⁻.

   $$Ag+_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$$

   0.623 moles Ag produced

   m = nm (molar mass) = (0.623 moles)(107.9 g/mol) = 67.2 g

   67.2 g of silver is deposited.

4. What products would be expected at each electrode? Write the half reactions and calculate the voltage required to electrolyze the following:

   a. aqueous copper (II) fluoride

      Species present: Cu^{2+}, F^{-} and H_2O

      At the cathode, either Cu^{2+} or H_2O will be reduced.

      $$Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)} \quad E° = +0.34V$$

      $$2\,H_2O_{(l)} + 2e^{-} \rightarrow 2\,OH^{-}_{(aq)} + H_{2(g)} \quad E° = -0.83V$$

      At anode, either F^{-} or H_2O will be oxidized.

      $$2\,F^{-}_{(aq)} \rightarrow F_{2(g)} + 2e^{-} \quad E° = -2.87V$$

      $$H_2O_{(l)} \rightarrow 2\,H+_{(aq)} + \tfrac{1}{2}O_{2(g)} + 2e^{-} \quad E° = -1.23V$$

      From the half-cell potentials, we can predict

      Cathode: copper

      Anode: oxygen gas and H^{+} ions

      $$E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (+0.34) = -0.89V$$

      At least 0.89 V is needed.

b. molten magnesium chloride

Cathode: $Mg^{2+}_{(aq)} + 2e^{-} \rightarrow Mg_{(s)}$ $E° = -2.37V$

Anode: $2\,Cl^{-}_{(aq)} \rightarrow Cl_{2(g)} + 2e^{-}$ $E° = -1.36V$

$E°_{cell} = E°_{ox} + E°_{red} = -1.36V + (-2.37V) = -3.73V$

At least 3.73V is needed.

c. aqueous iron (II) sulphate

Species: Fe^{2+}, SO_4^{2-}, H_2O

Cathode: $Fe^{2+}_{(aq)} + 2e^{-} \rightarrow Fe_{(s)}$ $E° = -0.44V$

Anode: $H_2O_{(l)} \rightarrow 2\,H^{+}_{(aq)} + \frac{1}{2}O_{2(g)} + 2e^{-}$ $E° = -1.23V$

$E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (-0.44V) = -1.67V$

Minimum voltage required is 1.67V.

d. molten silver bromide

Cathode: $Ag^{+}_{(l)} + 1e^{-} \rightarrow Ag_{(l)}$ $E° = +0.80V$

Anode: $2\,Br^{-}_{(l)} \rightarrow Br_{2(g)} + 2e^{-}$ $E° = -1.06V$

$E°_{cell} = E°_{ox} + E°_{red} = -1.06V + (+0.80V) = -0.26V$

At least 0.26V is needed.

e. aqueous aluminum chloride

Species: Al^{3+}, Cl^{-}, H_2O

Cathode: $2\,H_2O_{(l)} + 2e^{-} \rightarrow 2\,OH^{-}_{(aq)} + H_{2(g)}$ $E° = -0.83V$

Anode: $H_2O_{(l)} \rightarrow 2\,H+_{(aq)} + \frac{1}{2}O_{2(g)} + 2e^{-}$ $E° = -1.23V$

$E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (-0.83V) = -2.06V$

Minimum voltage required is 2.06V.

5. How many seconds would be needed to generate 3.00 moles of electrons from 10.0 amp of current?

$$\mathbf{t = \frac{n_{e^-} \times 96500\ C/mol}{I} = \frac{\left(3.00\ moles\ e^{-}\right)\left(96500\ C/mol\right)}{10.0\ A} = 28950\ s}$$

28950 s are needed.

GRADE 12 CHEMISTRY

Final Practice Examination

Name: ______________________________________

Student Number: ____________________________

Attending ☐ Non-Attending ☐

Phone Number: ______________________________

Address: ___________________________________

____

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For Marker's Use Only

Date: ______________________________

Final Mark _______ /100 = ________%

Comments:

Instructions

You have a maximum of 3 hours to complete your final examination.
The final examination will be weighted as follows:

- Modules 4 to 6: 100%

The format of the examination will be as follows:

- Part A: Multiple Choice *(23 x 1 = 23 marks)*
- Part B: Fill-in-the-Blanks *(19 x 1 = 19 marks)*
- Part C: Short Answer *(58 marks)*

Include units with all answers as required.

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- scrap paper
- a ruler
- a graphing or scientific calculator

The following resources are provided at the end of this examination:

- Periodic Table of Elements
- Alphabetical Listing of the Elements and Their Atomic Masses
- Table of Electronegativities
- Relative Strengths of Acids Table
- Solubility Chart
- Table of Standard Reduction Potentials with Values
- Names, Formulas, and Charges of Common Ions
- Common Ions

Part A: Multiple Choice *(23 marks total)*

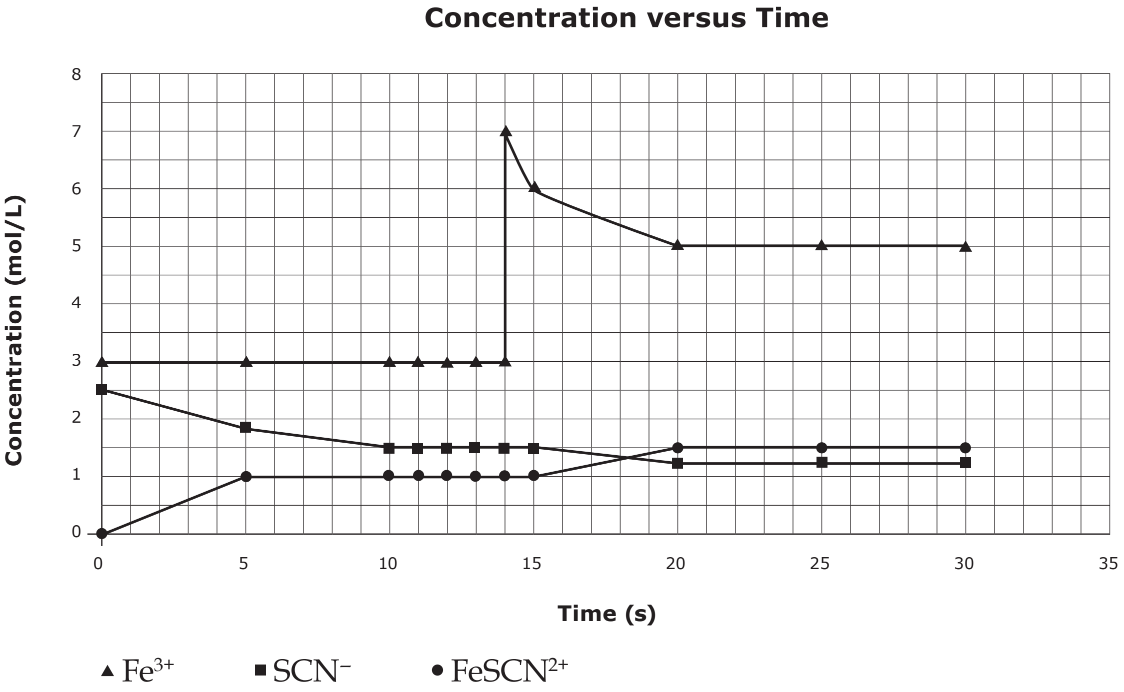
Use the answer sheet found at the end of this examination to answer the multiple-choice questions in this section. Shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Equilibrium *(7 marks)*

1. Based on the equilibrium constants given, in which of these reactions are the products most favoured over the reactants?
   a. $K_{eq} = 0.002$
   b. $K_{eq} = 0.0$
   c. $K_{eq} = 3.5$
   d. $K_{eq} = 6.0 \times 10^{-4}$
2. Which of the statements regarding the effect of reactant or product concentration changes on equilibrium is **false**?
   a. Adding a reactant always shifts a reversible reaction in the direction of reactants.
   b. Removing a reactant always shifts a reversible reaction in the direction of reactants.
   c. Adding a product causes a reversible reaction to shift in the direction of the formation of the reactants.
   d. Removing a product shifts a reversible reaction in the direction of formation of products.
3. Which of the following statements correctly describes the effect of an increase in temperature on the position of equilibrium for an exothermic reaction?
   a. The position of equilibrium does not change if you change the temperature.
   b. According to Le Châtelier's principle, the position of equilibrium will move to the left if the temperature is increased.
   c. If you increase the temperature, the position of equilibrium will move in such a way as to increase the temperature again.
   d. Equilibrium will shift to favour the reaction which releases heat in this context.
4. Which explanation best describes why a catalyst does **not** affect the position of equilibrium?
   a. A catalyst will increase the activation energy in both directions; therefore, no shift in the position of equilibrium will result.
   b. The position of equilibrium is not affected by adding a catalyst since a catalyst speeds up both the forward and back reactions by exactly the same amount.
   c. Equilibrium constants are not affected by adding a catalyst since no shift in the position of equilibrium occurs.
   d. Not all chemical reactions respond to catalysts.

Name: ____

The following graph represents the equilibrium established for the chemical reaction, $Fe^{3+} + SCN^{-} \leftrightarrow FeSCN^{2+}$



Use the graph for the next two questions.

5. At what time did this system first reach equilibrium?
   a. 5 s
   b. 10 s
   c. 15 s
   d. 20 s

6. What was the stress that occurred at 14 seconds?
   a. the addition of Fe^{3+}
   b. the addition of SCN^{-}
   c. the removal of $FeSCN^{2+}$
   d. the addition of a catalyst

7. Which of the following does not describe a use of the Haber process?
   a. the creation of nitrogen-containing explosives in WWI
   b. the production of ammonia for household cleaners such as Windex
   c. the creation of ammonia for industrial and household fertilizers
   d. the distillation of petroleum products

Acids and Bases *(9 marks)*

8. Which of the following statements **does not** support Arrhenius' work with acids and bases?
   a. Arrhenius discovered that all acidic and basic solutions he tested were electrolytes.
   b. He determined that acids and bases must ionize or dissociate in water.
   c. According to Arrhenius, an acid is defined as a hydrogen-containing compound that ionizes to yield hydrogen ions (H^+) dissolved in an aqueous solution.
   d. According to Arrhenius, an acid is defined as a compound that ionizes to yield hydroxide ions (OH^-) in an aqueous solution.

9. Which of the following is considered a limitation of the Brønsted-Lowry definition of acids and bases?
   a. The Brønsted-Lowry theory expands the definition of an acid and base to a proton donor or acceptor.
   b. Even though acids and bases can occur without water, there is still a requirement for the presence of a solvent.
   c. The Brønsted-Lowry theory cannot explain why some substances without hydroxides, like ammonia, can act as bases.
   d. The Brønsted-Lowry theory does not explain how substances without protons can act as acids.

10. Which of these descriptions is **not** a characteristic of the Lewis theory?
   a. This theory limits the number of compounds called acids, since any compound that has one or more valence shell orbitals cannot act as an acid.
   b. The Lewis definition is so general that any reaction in which a pair of electrons is transferred becomes an acid-base reaction.
   c. Lewis acid-base reactions include many reactions that would not be included with the Brønsted-Lowry definition.
   d. The Lewis acid-base theory does not affect Brønsted-Lowry bases because any Brønsted-Lowry base must have a pair of non-bonding electrons in order to accept a proton.

11. At 25°C, the concentrations of the hydronium and hydroxide ions in water are equal at
   a. 1.0×10^{7} mol/L
   b. 1.0×10^{14} mol/L
   c. 1.0×10^{-7} mol/L
   d. 1.0×10^{-14} mol/L

12. Which statement regarding indicators is **false**?
    a. They are strong, organic acids that change colour when the hydronium or hydroxide ion concentration is changed.
    b. Indicators change colour over a given pH range (the colour of each indicator can be coordinated with the pH).
    c. The colour of the indicator can be compared to a standard to determine the pH of the solution.
    d. Using an indicator is not as accurate as a pH metre.

13. All of these are examples of strong acids with the exception of
    a. $HClO_4$ (perchloric acid)
    b. HI (hydroiodic acid)
    c. HBr (hydrobromic acid)
    d. H_2CO_3 (carbonic acid)

14. A carefully controlled neutralization reaction is also known as a/an
    a. oxidation
    b. reduction
    c. titration
    d. dissociation

15. On a titration curve, what indicates when neutralization of the unknown has occurred?
    a. The beginning of the horizontal region in the acid pH range.
    b. The beginning of the horizontal region in the basic pH range.
    c. The point halfway on the vertical region of the graph.
    d. The beginning of the vertical region of the graph.

16. All of the following salts are BASIC in water with the exception of
    a. sodium nitrate ($NaNO_3$)
    b. sodium acetate ($NaC_2H_3O_2$)
    c. potassium phosphate (K_3PO_4)
    d. sodium sulphide (Na_2S)

Electrochemistry ***(7 marks)***

17. Using the Activity Series, determine which species is the strongest oxidizing agent.
    a. $F^-_{(aq)}$
    b. $F_{2\,(g)}$
    c. $Li^+_{(aq)}$
    d. $Li_{(s)}$

18. Which type of cell converts chemical energy into electrical energy from a spontaneous redox reaction?
    a. an electrolytic cell
    b. a half cell
    c. a wet cell
    d. a voltaic cell

19. Which statement about an anode in a voltaic cell is **false**?
    a. The anode is the electrode at which oxidation occurs.
    b. The anode is the electrode at which reduction occurs.
    c. Electrons are lost from this electrode.
    d. It is normally labelled the negative electrode.

20. Which of the following components is **not** part of a voltaic cell?
    a. a battery
    b. two half-reaction cells
    c. an external wire
    d. the salt bridge

21. Which of the following statements related to cell potential is **false**?
    a. The cell potential for a voltaic cell is a measure of the amount of voltage that can be generated by driving an electric current through a wire.
    b. One joule of energy is produced when one coulomb of electrical charge is transported across a potential of one volt.
    c. This difference in potential difference is an indication of how much energy is available to move electrons from the anode to the cathode.
    d. Electric charge can only flow between two electrodes when there is an equal amount of electric potential between the two points.

Name: ____

22. A short-hand notation is often used to represent the voltaic cell diagram. What does the double vertical line represent?
    a. It represents a phase boundary between the metal and the ion in solution.
    b. It represents the salt bridge.
    c. It represents the anode.
    d. It represents the cathode.

23. Many "reference" electrodes were tried before one was chosen as the standard to which all other electrodes would be measured. Which half cell was finally chosen as the reference?
    a. oxygen
    b. copper
    c. hydrogen
    d. zinc

Part B: Fill-in-the-Blank *(20 marks total)*

Use the answer sheet found at the end of this examination to answer the fill-in-the-blank questions of this section. Write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Using a term from the word bank provided below, complete each of the statements that follow. Each blank is worth one mark; therefore, some questions have a total value of two marks. There are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use.

acid
acidic
amphoteric
anode
Arrhenius
base
basic
brine
cathode
closed
closed system
concentration
constant
Daniell
Davy
dynamic
electrolysis
electrolyte
electrons
electroplating
energy
equal
equilibrium
equilibrium constant
equivalence
exothermix
forward
Galvani
heterogeneous
homogeneous
hydrolysis
hydronium
increases
indicator(s)
K_c
K_{eq}
Lavoisier
Le Châtelier's principle
Lewis
Liebig
mass action
neutral
oxidation
oxidizing
pH
pOH
pressure
proton
reducing
reduction
reversible
salt bridge
spontaneous
standard reduction potentials
strong
temperature
titration
Volta

Name: ____

Equilibrium ***(6 marks)***

1. Equilibrium can only be established when all particles are kept in a sealed container and certain conditions are kept _______________.
2. In a reversible reaction, when the rate at which products are formed equals the rate at which reactants are formed, we say that the reaction has reached a _______________ equilibrium.
3. Equilibrium constants are specific for only one reaction at a particular _____________.
4. Le Châtelier's principle states that the equilibrium of a _______________ _______________ will shift so as to relieve the stress, thereby restoring equilibrium.
5. If the forward reaction is endothermic, increasing the temperature _______________ the value of the equilibrium constant.

Acids and Bases ***(8 marks)***

6. _______________ studied the conductivity of solutions and proposed that electrolytes break up into charged particles in water.
7. When water acts as the base, accepting the proton, the result is the H_3O^+ ion called the _______________ ion.
8. A conjugate base is what remains after a/an _______________ has donated its proton.
9. A neutral solution occurs when the hydronium ion concentration is _______________ to the hydroxide ion concentration.
10. In general, a strong acid will dissociate close to 100% and have a very large _______________.
11. An _______________ is defined as a compound that conducts an electric current when it is in an aqueous state.
12. At equivalence, the _______________ ion and hydroxide ion concentrations are equal.
13. _______________ salts result from the reaction of a strong base with a weak acid.

Electrochemistry ***(6 marks)***

14. A spontaneous reaction is one that occurs without any added _______________.
15. The stronger the oxidizing agent, the weaker the _______________ ability.
16. In an electrochemical cell, the electrode at which reduction occurs is called the _______________.
17. _______________ flow in the external circuit from the anode to the cathode of an electrochemical cell.
18. All the half reactions on the Table of Standard Reduction Potentials with Values are written as _______________.
19. _______________ is a process where an electric current is used to plate a metal onto another surface.

Part C: Short Answer *(57 marks total)*

Answer each of the following questions 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 as well as the appropriate unit.

Equilibrium *(13 marks)*

1. Explain how reaction rate and equilibrium are related concepts. Give an example to illustrate this relationship. *(2 marks)*

2. Write the equilibrium law (mass action expression) for the following reaction. *(2 marks)*

   $2\ SO_{2(g)} + O_{2(g)} \leftrightarrow 2\ SO_{3(g)}$

Name: ____

3. Explain what an equilibrium constant of 1 indicates about the reaction. *(2 marks)*

4. For the reaction, $\mathbf{CO_{(g)} + 2\ H_{2(g)} \leftrightarrow CH_3OH_{(g)}}$ **+ energy**, predict the effect of the following changes on the equilibrium concentration of $CH_3OH_{(g)}$:
   a. an increase in temperature *(1 mark)*

   b. a decrease in pressure *(1 mark)*

   c. the addition of $CO_{(g)}$ *(1 mark)*

5. How might Le Châtelier’s principle be useful in the chemical industry? For example, how could you ensure a high yield in the production of ammonia? *(2 marks)*

6. Several types of stress can disrupt chemical equilibrium. Name two such stresses. *(2 marks)*

Name: ____

Acids and Bases *(23 marks)*

7. Identify the acid, base, conjugate acid, and conjugate base in the following reversible reaction. *(2 marks)*

   $HF_{(aq)} + HSO_3^{-}{}_{(aq)} \leftrightarrow F^{-}{}_{(aq)} + H_2SO_{3(aq)}$

8. Using the ionization of water equation given, predict the effect of dissolving a base on hydronium and hydroxide ion concentrations by using Le Châtelier's principle. *(3 marks)*

   $H_2O_{(l)} + H_2O_{(l)} \leftrightarrow H_3O^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$

9. What is the $[H_3O^+]$ in a solution with $[OH^-]$ of 5.67×10^{-3}? *(2 marks)*

10. In a neutral solution, the $[OH^-] = 1 \times 10^{-7}$ mol/L. Calculate the pOH of a neutral solution, showing all the steps of your work. *(2 marks)*

11. Are all strong acids also strong electrolytes? Explain your answer. *(2 marks)*

Name: ____

12. Calculate the percent dissociation of a 0.20 mol/L solution of the weak acid, HNO_2, if the pH of the solution is 4.20. The dissociation reaction has been provided. Show all the steps of your work. *(5 marks)*

    $HNO_{2(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_2^-{}_{(aq)}$

13. Calculate the concentration of the acid, if 25.0 mL of H_3PO_4 is required to neutralize 29.0 mL of 0.830 mol/L KOH. Complete the neutralization reaction and then show all the steps of your work. *(7 marks)*

$H_3PO_{4(aq)} + 3\ KOH_{(aq)} \rightarrow$

Name: ____

Electrochemistry ***(21 marks)***

14. Explain how the activity series can be used to predict if a reaction will be spontaneous or not. *(2 marks)*

15. Two half cells are connected under standard conditions to make an electrochemical cell. The two half cells are a copper-copper(I) ion (Cu/Cu^+) and an aluminum-aluminum ion (Al/Al^{3+}). Using your Table of Standard Reduction Potentials with Values complete the following:

    a. Write the two half reactions that will occur within this electrochemical cell. *(2 marks)*

    b. Identify which reaction is the anode and which is the cathode. *(0.5 mark x 2 = 1 mark)*

c. Calculate the overall voltage of the cell. *(2 marks)*

d. Write the overall balanced redox reaction for the electrochemical cell. *(3 marks)*

Name: ____

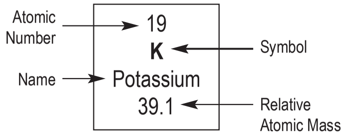
e. Draw a labelled electrochemical cell indicating
  - the electrodes (anode, cathode, positive, and negative) *(2 marks)*
  - the types of ions *(2 marks)*
  - the direction of electron flow *(0.5 mark)*
  - the electrode half reactions *(2 marks)*
  - the salt bridge *(0.5 mark)*

16. Calculate how long it would take an aqueous gold (III) chloride cell to plate 2.5 g of gold on a bracelet using a current of 2.5 A. The half reaction has been provided for you. *(4 marks)*

Half reaction: $Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

Name: ________________________________

Periodic Table of Elements



	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)

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)	Irdium	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)	Sulphur	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
Einstienium	(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		

Name: ____

Table of Electronegativities

	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** —	1
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** —	2
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** —	3
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** —	4
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** —	5
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** —	6
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** —	7

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** —

Relative Strengths of Acids Table

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulphuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4×10^{-2}
Sulphurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7×10^{-2}
Hydrogen sulphate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3×10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1×10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0×10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3×10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7×10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1×10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7×10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5×10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6×10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4×10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4×10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4×10^{-7}
Hydrogen sulphide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3×10^{-8}
Hydrogen sulphite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2×10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7×10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7×10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0×10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4×10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4×10^{-13}
Hydrogen sulphide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2×10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8×10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Name: ____

Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	low solubility
	all others	soluble
sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulphide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulphite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

Table of Standard Reduction Potentials with Values

Oxidized species	↔	Reduced Species	E°/V
$Li^{+}{}_{(aq)} + e^{-}$	↔	$Li_{(s)}$	−3.04
$K^{+}{}_{(aq)} + e^{-}$	↔	$K_{(s)}$	−2.93
$Ca^{2+}{}_{(aq)} + 2e^{-}$	↔	$Ca_{(s)}$	−2.87
$Na^{+}{}_{(aq)} + e^{-}$	↔	$Na_{(s)}$	−2.71
$Mg^{2+}{}_{(aq)} + 2e^{-}$	↔	$Mg_{(s)}$	−2.37
$Al^{3+}{}_{(aq)} + 3e^{-}$	↔	$Al_{(s)}$	−1.66
$Mn^{2+}{}_{(aq)} + 2e^{-}$	↔	$Mn_{(s)}$	−1.19
$H_2O_{(l)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)} + OH^{-}{}_{(aq)}$	−0.83
$Zn^{2+}{}_{(aq)} + 2e^{-}$	↔	$Zn_{(s)}$	−0.76
$Fe^{2+}{}_{(aq)} + 2e^{-}$	↔	$Fe_{(s)}$	−0.45
$Ni^{2+}{}_{(aq)} + 2e^{-}$	↔	$Ni_{(s)}$	−0.26
$Sn^{2+}{}_{(aq)} + 2e^{-}$	↔	$Sn_{(s)}$	−0.14
$Pb^{2+}{}_{(aq)} + 2e^{-}$	↔	$Pb_{(s)}$	−0.13
$H^{+}{}_{(aq)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}{}_{(aq)} + e^{-}$	↔	$Cu^{+}{}_{(aq)}$	+0.15
$SO_4{}^{2-}{}_{(aq)} + 4\ H^{+}{}_{(aq)} + 2e^{-}$	↔	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}{}_{(aq)} + 2e^{-}$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^{-}$		$2OH^{-}{}_{(aq)}$	+0.40
$Cu^{+}{}_{(aq)} + e^{-}$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^{-}$	↔	$I^{-}{}_{(aq)}$	+0.54
$Fe^{3+}{}_{(aq)} + e^{-}$	↔	$Fe^{2+}{}_{(aq)}$	+0.77
$Ag^{+}{}_{(aq)} + e^{-}$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^{-}$	↔	$Br^{-}{}_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^{+}{}_{(aq)} + 2e^{-}$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7{}^{2-}{}_{(aq)} + 14H^{+}{}_{(aq)} + 6e^{-}$	↔	$2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^{-}$	↔	$Cl^{-}{}_{(aq)}$	+1.36
$MnO_4{}^{-}{}_{(aq)} + 8H^{+}{}_{(aq)} + 5e^{-}$	↔	$Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^{-}$	↔	$F^{-}{}_{(aq)}$	+2.87

Name: ____

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

Names, Formulas, and Charges of Common Ions *(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^-	sulphate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulphate	HSO_4^-
fluoride	F^-	sulphide	S^{2-}
hydroxide	OH^-	hydrogen sulphide	HS^-
hypochlorite	ClO^-	sulphite	SO_3^{2-}
iodide	I^-	hydrogen sulphite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Name: ____

Common Ions

Cations (Positive Ions)

1^{+} charge		2^{+} charge		3^{+} charge	
$NH4^{+}$	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	$Pb2^{2+}$	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

Common Ions *(continued)*

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulphide	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-}	Sulphate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulphide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulphite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulphate
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 sulphate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Name: ____

Final Practice Examination Answer Sheet

Part A: Multiple Choice *(23 marks)*

For each multiple-choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Example: Ⓐ Ⓑ ● Ⓓ

1. Ⓐ Ⓑ Ⓒ Ⓓ
2. Ⓐ Ⓑ Ⓒ Ⓓ
3. Ⓐ Ⓑ Ⓒ Ⓓ
4. Ⓐ Ⓑ Ⓒ Ⓓ
5. Ⓐ Ⓑ Ⓒ Ⓓ
6. Ⓐ Ⓑ Ⓒ Ⓓ
7. Ⓐ Ⓑ Ⓒ Ⓓ
8. Ⓐ Ⓑ Ⓒ Ⓓ
9. Ⓐ Ⓑ Ⓒ Ⓓ
10. Ⓐ Ⓑ Ⓒ Ⓓ
11. Ⓐ Ⓑ Ⓒ Ⓓ
12. Ⓐ Ⓑ Ⓒ Ⓓ
13. Ⓐ Ⓑ Ⓒ Ⓓ
14. Ⓐ Ⓑ Ⓒ Ⓓ
15. Ⓐ Ⓑ Ⓒ Ⓓ
16. Ⓐ Ⓑ Ⓒ Ⓓ
17. Ⓐ Ⓑ Ⓒ Ⓓ
18. Ⓐ Ⓑ Ⓒ Ⓓ
19. Ⓐ Ⓑ Ⓒ Ⓓ
20. Ⓐ Ⓑ Ⓒ Ⓓ
21. Ⓐ Ⓑ Ⓒ Ⓓ
22. Ⓐ Ⓑ Ⓒ Ⓓ
23. Ⓐ Ⓑ Ⓒ Ⓓ

Part B: Fill-in-the-Blank *(20 marks)*

For each fill-in-the-blank question, write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Equilibrium *(6 marks)*

1. ____
2. ____
3. ____
4. ____

   ____
5. ____

continued

Acids and Bases *(8 marks)*

6. ____
7. ____
8. ____
9. ____
10. ____
11. ____
12. ____
13. ____

Electrochemistry *(6 marks)*

14. ____
15. ____
16. ____
17. ____
18. ____
19. ____

GRADE 12 CHEMISTRY

Final Practice Examination
Answer Key

Name: ____

Student Number: ______________________________

Attending ☐ Non-Attending ☐

Phone Number: ________________________________

Address: _____________________________________

____

____

For Marker's Use Only

Date: ________________________

Final Mark ______ /100 = ________%

Comments:

Instructions

You have a maximum of 3 hours to complete your final examination.
The final examination will be weighted as follows:

- Modules 4 to 6: 100%

The format of the examination will be as follows:

- Part A: Multiple Choice *(23 x 1 = 23 marks)*
- Part B: Fill-in-the-Blanks *(19 x 1 = 19 marks)*
- Part C: Short Answer *(58 marks)*

Include units with all answers as required.

You will need the following in order to complete this examination:

- writing utensils and eraser or correction fluid
- scrap paper
- a ruler
- a graphing or scientific calculator

The following resources are provided at the end of this examination:

- Periodic Table of Elements
- Alphabetical Listing of the Elements and Their Atomic Masses
- Table of Electronegativities
- Relative Strengths of Acids Table
- Solubility Chart
- Table of Standard Reduction Potentials with Values
- Names, Formulas, and Charges of Common Ions
- Common Ions

Part A: Multiple Choice *(23 marks total)*

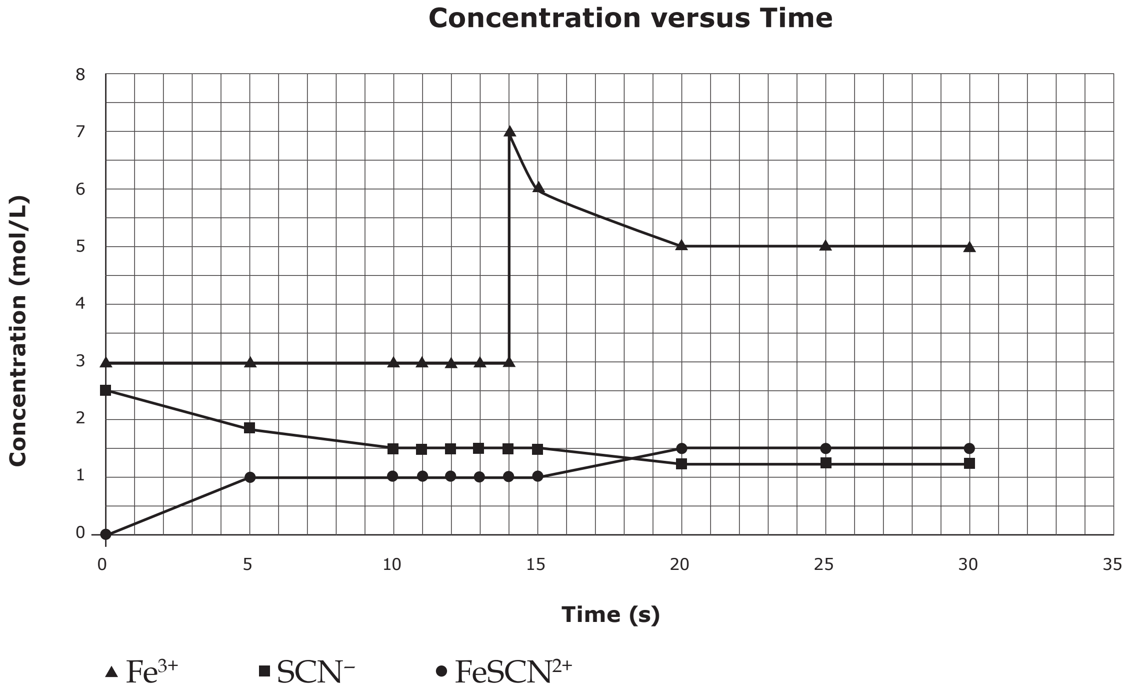
Use the answer sheet found at the end of this examination to answer the multiple-choice questions in this section. Shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Equilibrium *(7 marks)*

1. Based on the equilibrium constants given, in which of these reactions are the products most favoured over the reactants?
   a. $K_{eq} = 0.002$
   b. $K_{eq} = 0.0$
   c. $K_{eq} = 3.5$
   d. $K_{eq} = 6.0 \times 10^{-4}$
2. Which of the statements regarding the effect of reactant or product concentration changes on equilibrium is **false**?
   a. Adding a reactant always shifts a reversible reaction in the direction of reactants.
   b. Removing a reactant always shifts a reversible reaction in the direction of reactants.
   c. Adding a product causes a reversible reaction to shift in the direction of the formation of the reactants.
   d. Removing a product shifts a reversible reaction in the direction of formation of products.
3. Which of the following statements correctly describes the effect of an increase in temperature on the position of equilibrium for an exothermic reaction?
   a. The position of equilibrium does not change if you change the temperature.
   b. According to Le Châtelier’s principle, the position of equilibrium will move to the left if the temperature is increased.
   c. If you increase the temperature, the position of equilibrium will move in such a way as to increase the temperature again.
   d. Equilibrium will shift to favour the reaction which releases heat in this context.
4. Which explanation best describes why a catalyst does **not** affect the position of equilibrium?
   a. A catalyst will increase the activation energy in both directions; therefore, no shift in the position of equilibrium will result.
   b. The position of equilibrium is not affected by adding a catalyst since a catalyst speeds up both the forward and back reactions by exactly the same amount.
   c. Equilibrium constants are not affected by adding a catalyst since no shift in the position of equilibrium occurs.
   d. Not all chemical reactions respond to catalysts.

Name: ____

The following graph represents the equilibrium established for the chemical reaction, $Fe^{3+} + SCN^{-} \leftrightarrow FeSCN^{2+}$



Use the graph for the next two questions.

5. At what time did this system first reach equilibrium?
   - a. 5 s
   - b. 10 s
   - c. 15 s
   - d. 20 s
6. What was the stress that occurred at 14 seconds?
   - a. the addition of Fe^{3+}
   - b. the addition of SCN^{-}
   - c. the removal of $FeSCN^{2+}$
   - d. the addition of a catalyst
7. Which of the following does not describe a use of the Haber process?
   - a. the creation of nitrogen-containing explosives in WWI
   - b. the production of ammonia for household cleaners such as Windex
   - c. the creation of ammonia for industrial and household fertilizers
   - d. the distillation of petroleum products

Acids and Bases ***(9 marks)***

8. Which of the following statements **does not** support Arrhenius' work with acids and bases?
   a. Arrhenius discovered that all acidic and basic solutions he tested were electrolytes.
   b. He determined that acids and bases must ionize or dissociate in water.
   c. According to Arrhenius, an acid is defined as a hydrogen-containing compound that ionizes to yield hydrogen ions (H^+) dissolved in an aqueous solution.
   d. According to Arrhenius, an acid is defined as a compound that ionizes to yield hydroxide ions (OH^-) in an aqueous solution.

9. Which of the following is considered a limitation of the Brønsted-Lowry definition of acids and bases?
   a. The Brønsted-Lowry theory expands the definition of an acid and base to a proton donor or acceptor.
   b. Even though acids and bases can occur without water, there is still a requirement for the presence of a solvent.
   c. The Brønsted-Lowry theory cannot explain why some substances without hydroxides, like ammonia, can act as bases.
   d. The Brønsted-Lowry theory does not explain how substances without protons can act as acids.

10. Which of these descriptions is **not** a characteristic of the Lewis theory?
    a. This theory limits the number of compounds called acids, since any compound that has one or more valence shell orbitals cannot act as an acid.
    b. The Lewis definition is so general that any reaction in which a pair of electrons is transferred becomes an acid-base reaction.
    c. Lewis acid-base reactions include many reactions that would not be included with the Brønsted-Lowry definition.
    d. The Lewis acid-base theory does not affect Brønsted-Lowry bases because any Brønsted-Lowry base must have a pair of non-bonding electrons in order to accept a proton.

11. At 25°C, the concentrations of the hydronium and hydroxide ions in water are equal at
    a. 1.0×10^{7} mol/L
    b. 1.0×10^{14} mol/L
    c. 1.0×10^{-7} mol/L
    d. 1.0×10^{-14} mol/L

Name: ____

12. Which statement regarding indicators is **false**?
    a. They are strong, organic acids that change colour when the hydronium or hydroxide ion concentration is changed.
    b. Indicators change colour over a given pH range (the colour of each indicator can be coordinated with the pH).
    c. The colour of the indicator can be compared to a standard to determine the pH of the solution.
    d. Using an indicator is not as accurate as a pH metre.

13. All of these are examples of strong acids with the exception of
    a. $HClO_4$ (perchloric acid)
    b. HI (hydroiodic acid)
    c. HBr (hydrobromic acid)
    d. H_2CO_3 (carbonic acid)

14. A carefully controlled neutralization reaction is also known as a/an
    a. oxidation
    b. reduction
    c. titration
    d. dissociation

15. On a titration curve, what indicates when neutralization of the unknown has occurred?
    a. The beginning of the horizontal region in the acid pH range.
    b. The beginning of the horizontal region in the basic pH range.
    c. The point halfway on the vertical region of the graph.
    d. The beginning of the vertical region of the graph.

16. All of the following salts are BASIC in water with the exception of
    a. sodium nitrate ($NaNO_3$)
    b. sodium acetate ($NaC_2H_3O_2$)
    c. potassium phosphate (K_3PO_4)
    d. sodium sulphide (Na_2S)

Electrochemistry ***(7 marks)***

17. Using the Activity Series, determine which species is the strongest oxidizing agent.
    a. $F^-_{(aq)}$
    b. $F_{2\,(g)}$
    c. $Li^+_{(aq)}$
    d. $Li_{(s)}$

18. Which type of cell converts chemical energy into electrical energy from a spontaneous redox reaction?
    a. an electrolytic cell
    b. a half cell
    c. a wet cell
    d. a voltaic cell

19. Which statement about an anode in a voltaic cell is **false**?
    a. The anode is the electrode at which oxidation occurs.
    b. The anode is the electrode at which reduction occurs.
    c. Electrons are lost from this electrode.
    d. It is normally labelled the negative electrode.

20. Which of the following components is **not** part of a voltaic cell?
    a. a battery
    b. two half-reaction cells
    c. an external wire
    d. the salt bridge

21. Which of the following statements related to cell potential is **false**?
    a. The cell potential for a voltaic cell is a measure of the amount of voltage that can be generated by driving an electric current through a wire.
    b. One joule of energy is produced when one coulomb of electrical charge is transported across a potential of one volt.
    c. This difference in potential difference is an indication of how much energy is available to move electrons from the anode to the cathode.
    d. Electric charge can only flow between two electrodes when there is an equal amount of electric potential between the two points.

Name: ____________________________________

22. A short-hand notation is often used to represent the voltaic cell diagram. What does the double vertical line represent?
    a. It represents a phase boundary between the metal and the ion in solution.
    b. It represents the salt bridge.
    c. It represents the anode.
    d. It represents the cathode.

23. Many “reference” electrodes were tried before one was chosen as the standard to which all other electrodes would be measured. Which half cell was finally chosen as the reference?
    a. oxygen
    b. copper
    c. hydrogen
    d. zinc

Part B: Fill-in-the-Blank *(20 marks total)*

Use the answer sheet found at the end of this examination to answer the fill-in-the-blank questions of this section. Write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Using a term from the word bank provided below, complete each of the statements that follow. Each blank is worth one mark; therefore, some questions have a total value of two marks. There are MORE terms provided than you need, so read over the list carefully and choose the terms you want to use.

acid
acidic
amphoteric
anode
Arrhenius
base
basic
brine
cathode
closed
closed system
concentration
constant
Daniell
Davy
dynamic
electrolysis
electrolyte
electrons
electroplating
energy
equal
equilibrium
equilibrium constant
equivalence
exothermix
forward
Galvani
heterogeneous
homogeneous
hydrolysis
hydronium
increases
indicator(s)
K_c
K_{eq}
Lavoisier
Le Châtelier's principle
Lewis
Liebig
mass action
neutral
oxidation
oxidizing
pH
pOH
pressure
proton
reducing
reduction
reversible
salt bridge
spontaneous
standard reduction potentials
strong
temperature
titration
Volta

Name: _______________________________________

Equilibrium *(6 marks)*

1. Equilibrium can only be established when all particles are kept in a sealed container and certain conditions are kept _______________.
2. In a reversible reaction, when the rate at which products are formed equals the rate at which reactants are formed, we say that the reaction has reached a _______________ equilibrium.
3. Equilibrium constants are specific for only one reaction at a particular _____________.
4. Le Châtelier's principle states that the equilibrium of a _______________ _______________ will shift so as to relieve the stress, thereby restoring equilibrium.
5. If the forward reaction is endothermic, increasing the temperature _______________ the value of the equilibrium constant.

Acids and Bases *(8 marks)*

6. _______________ studied the conductivity of solutions and proposed that electrolytes break up into charged particles in water.
7. When water acts as the base, accepting the proton, the result is the H_3O^+ ion called the _______________ ion.
8. A conjugate base is what remains after a/an _______________ has donated its proton.
9. A neutral solution occurs when the hydronium ion concentration is _______________ to the hydroxide ion concentration.
10. In general, a strong acid will dissociate close to 100% and have a very large _______________.
11. An _______________ is defined as a compound that conducts an electric current when it is in an aqueous state.
12. At equivalence, the _______________ ion and hydroxide ion concentrations are equal.
13. _______________ salts result from the reaction of a strong base with a weak acid.

Electrochemistry *(6 marks)*

14. A spontaneous reaction is one that occurs without any added _______________.
15. The stronger the oxidizing agent, the weaker the _______________ ability.
16. In an electrochemical cell, the electrode at which reduction occurs is called the _______________.
17. _______________ flow in the external circuit from the anode to the cathode of an electrochemical cell.
18. All the half reactions on the Table of Standard Reduction Potentials with Values are written as _______________.
19. _______________ is a process where an electric current is used to plate a metal onto another surface.

Part C: Short Answer *(57 marks total)*

Answer each of the following questions 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 as well as the appropriate unit.

Equilibrium *(13 marks)*

1. Explain how reaction rate and equilibrium are related concepts. Give an example to illustrate this relationship. *(2 marks)*

   When the rates of the forward and the reverse reactions are equal, equilibrium has been established. *(1 mark)* **For example, when a liquid is placed in a closed container, equilibrium is established when the liquid vaporizes and condenses at the same rate.** *(1 mark)*

2. Write the equilibrium law (mass action expression) for the following reaction. *(2 marks)*

   $2\ SO_{2(g)} + O_{2(g)} \leftrightarrow 2\ SO_{3(g)}$

   $$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

Name: ____

3. Explain what an equilibrium constant of 1 indicates about the reaction. *(2 marks)*

   If *K* **= 1, the ratio of [products] = [reactants] at equilibrium.** *(1 mark)*

   Neither reactants nor products are favoured at equilibrium. *(1 mark)*

4. For the reaction, **$CO_{(g)} + 2\ H_{2(g)} \leftrightarrow CH_3OH_{(g)}$ + energy**, predict the effect of the following changes on the equilibrium concentration of $CH_3OH_{(g)}$:

   a. an increase in temperature *(1 mark)*

   An increase in temperature decreases [CH_3OH].

   b. a decrease in pressure *(1 mark)*

   A decrease in pressure decreases [CH_3OH].

   c. the addition of $CO_{(g)}$ *(1 mark)*

   An addition of $CO_{(g)}$ increases [CH_3OH].

5. How might Le Châtelier’s principle be useful in the chemical industry? For example, how could you ensure a high yield in the production of ammonia? *(2 marks)*

   Le Châtelier’s principle could be used to maximize the amount of product formed in chemical reactions. *(1 mark)* **By maintaining high temperature and pressure, by constantly adding reactants, and by immediately removing products, the products continue to be favoured.** *(1 mark)*

6. Several types of stress can disrupt chemical equilibrium. Name two such stresses. *(2 marks)*

   Any two of the following for 1 mark each:

   - **changes in the concentration of reactants or products**
   - **temperature changes**
   - **pressure changes**
   - **volume changes**
   - **the addition of a catalyst**

Name: ____

Acids and Bases ***(23 marks)***

7. Identify the acid, base, conjugate acid, and conjugate base in the following reversible reaction. *(2 marks)*

   $HF_{(aq)} + HSO_3^-{}_{(aq)} \leftrightarrow F^-{}_{(aq)} + H_2SO_{3(aq)}$

   acid: $HF_{(aq)}$
   base: $HSO_3^-{}_{(aq)}$
   conjugate acid: $H_2SO_{3(aq)}$
   conjugate base: $F^-{}_{(aq)}$

8. Using the ionization of water equation given, predict the effect of dissolving a base on hydronium and hydroxide ion concentrations by using Le Châtelier's principle. *(3 marks)*

   $H_2O_{(l)} + H_2O_{(l)} \leftrightarrow H_3O^+{}_{(aq)} + OH^-{}_{(aq)}$

   When a base is dissolved in water, the hydroxide ion concentration increases. *(1 mark)* **According to Le Châtelier's principle, the equilibrium shifts left to use up some of the added hydroxide and maintain K_w at 1.0×10^{-14}.** *(1 mark)* **Since equilibrium shifts left, the hydronium ion concentration is reduced.** *(1 mark)* **As such, adding a base to water increases the hydroxide ion concentration and reduces the hydronium ion concentration.**

9. What is the [H_3O^+] in a solution with [OH^-] of 5.67×10^{-3}? *(2 marks)*

$$[H_3O^+] = \frac{K_W}{[OH^-]} = \frac{1.0\times10^{-14}}{5.67\times10^{-3}} = 1.76\times10^{-12}\ \text{mol/L } H_3O^+$$

10. In a neutral solution, the [OH^-] = 1×10^{-7} mol/L. Calculate the pOH of a neutral solution, showing all the steps of your work. *(2 marks)*

pOH = –log [OH^-] *(1 mark)*

pOH = –(log OH^-) = (–log 10^{-7}) = 7.0 *(1 mark)*

11. Are all strong acids also strong electrolytes? Explain your answer. *(2 marks)*

Yes, all strong acids are also by definition strong electrolytes. In water, strong electrolytes completely dissociate into ions, and become good conductors of electricity.

Name: ____

12. Calculate the percent dissociation of a 0.20 mol/L solution of the weak acid, HNO_2, if the pH of the solution is 4.20. The dissociation reaction has been provided. Show all the steps of your work. *(5 marks)*

$HNO_{2(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_2^-{}_{(aq)}$

$\mathbf{[H_3O^+] = 10^{-pH} = 10^{-4.20} = 6.3 \times 10^{-5}\ mol/L}$ *(2 marks)*

$$\textbf{\% dissociation} = \frac{[H_3O^+]}{[HNO_2]} \times 100 = \frac{6.3 \times 10^{-5}\ \text{mol/L}}{0.2\ \text{mol/L}} \times 100 \quad \textit{(2 marks)}$$

% dissociation = 0.0315 % *(1 mark)*

13. Calculate the concentration of the acid, if 25.0 mL of H_3PO_4 is required to neutralize 29.0 mL of 0.830 mol/L KOH. Complete the neutralization reaction and then show all the steps of your work. *(7 marks)*

$H_3PO_{4(aq)} + 3\ KOH_{(aq)} \rightarrow$

$\mathbf{H_3PO_{4(aq)} + 3\ KOH_{(aq)} \rightarrow K_3PO_{4(aq)} + 3\ H_2O_{(l)}}$ *(1 mark)*

moles of KOH = C x V = (0.830 mol/L)(0.0290 L) = 0.02407 moles KOH *(2 marks)*

$$\mathbf{moles\ H_3PO_4 = 0.02407\ moles\ KOH\left(\frac{1\ mole\ H_3PO_4}{3\ moles\ KOH}\right) = 0.008023\ moles\ H_3PO_4}$$ *(2 marks)*

$$\mathbf{Concentration = \frac{mol}{V} = \frac{0.008023\ moles}{0.0250\ L} = 0.321\ mol/L}$$ *(2 marks)*

The student may have constructed the following cvn table:

	3 KOH	H_3PO_4
C (mol/L)	0.830 mol/L	0.321 mol/L
V (L)	0.029 L	0.025 L
n (mol)	0.02407 mol	0.008023 mol

Name: ____

Electrochemistry *(21 marks)*

14. Explain how the activity series can be used to predict if a reaction will be spontaneous or not. *(2 marks)*

    The spontaneity rule states that any species on the left side of the activity series will react spontaneously and oxidize any species on the right side which is above it. In other words, "spontaneous reactions are up and to the right."

15. Two half cells are connected under standard conditions to make an electrochemical cell. The two half cells are a copper-copper(I) ion (Cu/Cu^+) and an aluminum-aluminum ion (Al/Al^{3+}). Using your Table of Standard Reduction Potentials with Values, complete the following:

    a. Write the two half reactions that will occur within this electrochemical cell. *(2 marks)*

    1 mark for each half reaction, even if is not labelled

    $Cu^+ + e^- \rightarrow Cu_{(s)}$

    $Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$

    b. Identify which reaction is the anode and which is the cathode. *(0.5 mark x 2 = 1 mark)*

    Cathode: $Cu^+ + e^- \rightarrow Cu_{(s)}$ *(0.5 mark)*

    Anode: $Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$ *(0.5 mark)*

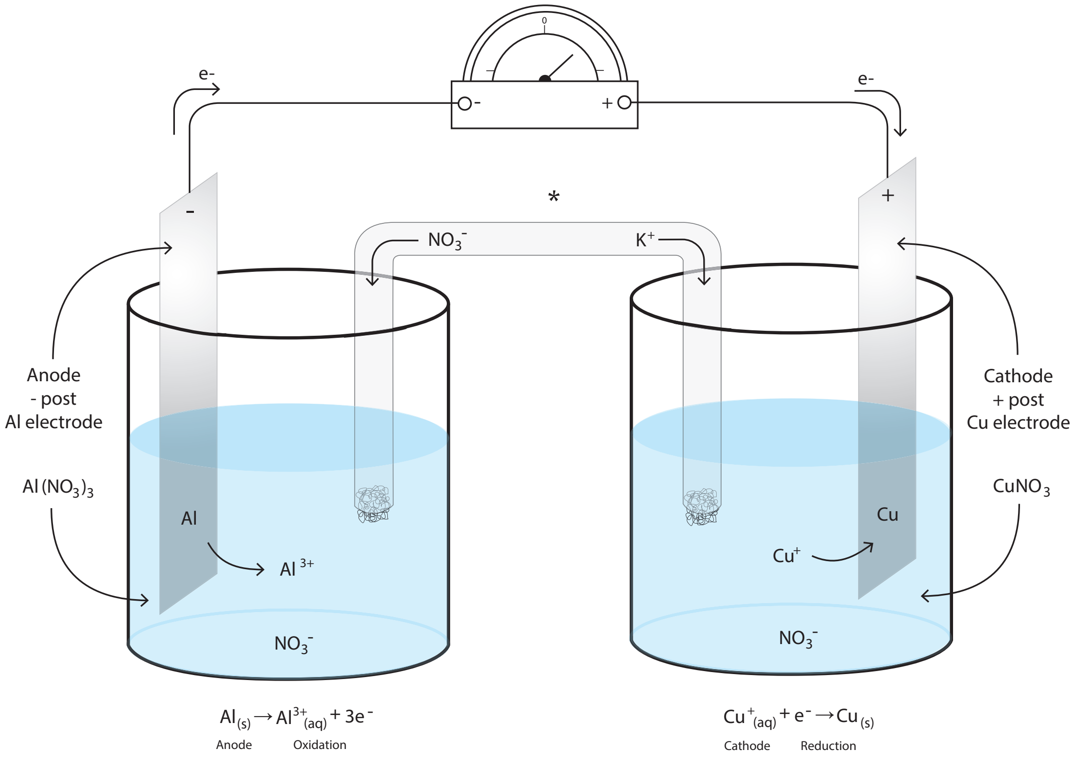
c. Calculate the overall voltage of the cell. *(2 marks)*

+0.52V + 1.66V = +2.18V

d. Write the overall balanced redox reaction for the electrochemical cell. *(3 marks)*

$$\begin{array}{l} (\times 3)\ \mathbf{Cu^{+} + e^{-} \rightarrow Cu_{(s)}} \\ \qquad \mathbf{Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^{-}} \\ \hline \qquad \mathbf{Al_{(s)} + 3\,Cu^{+}_{(aq)} \rightarrow Al^{3+}_{(aq)} + 3\,Cu_{(s)}} \end{array}$$

e. Draw a labelled electrochemical cell indicating
- the electrodes (anode, cathode, positive, and negative) *(2 marks)*
- the types of ions *(2 marks)*
- the direction of electron flow *(0.5 mark)*
- the electrode half reactions *(2 marks)*
- the salt bridge *(0.5 mark)*



*Note: Any combination of cation and anion is acceptable in the salt bridge.

16. Calculate how long it would take an aqueous gold (III) chloride cell to plate 2.5 g of gold on a bracelet using a current of 2.5 A. The half reaction has been provided for you. *(4 marks)*

Half reaction: $Au^{3+}_{(aq)} + 3e^{-} \rightarrow Au_{(s)}$

$$\textbf{moles Au} = \frac{m}{M} = \frac{2.50\ g}{197.0\ g/mol} = 0.01269\ \textbf{moles Au}$$

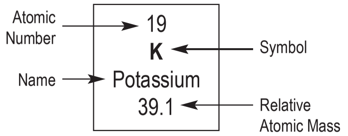
$$\textbf{moles } e^{-} = 0.01269\ \textbf{moles Au}\left(\frac{3\ \textbf{moles } e^{-}}{1\ \textbf{mole Zn}}\right) = 0.03807\ \textbf{moles } e^{-}$$

$$t = \frac{(96500\ C/mol)\,n_{e^{-}}}{I} = \frac{(96500\ C/mol)(0.03807\ \textbf{moles } e^{-})}{2.5\ A} = 1469.5\ s$$

It would require about 1470 s (24.5 minutes) to plate 2.5 g of gold.

Name: ____

Periodic Table of Elements



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** Unnhexium (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)

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)	Irdium	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)	Sulphur	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
Einstienium	(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		

Name: ____________________________________

Table of Electronegativities

	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** —	1
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** —	2
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** —	3
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** —	4
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** —	5
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** —	6
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** —	7

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** —

Relative Strengths of Acids Table

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulphuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4 x 10^{-2}
Sulphurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7 x 10^{-2}
Hydrogen sulphate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3 x 10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1 x 10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0 x 10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3 x 10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7 x 10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1 x 10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7 x 10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5 x 10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6 x 10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4 x 10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8 x 10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4 x 10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4 x 10^{-7}
Hydrogen sulphide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0 x 10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3 x 10^{-8}
Hydrogen sulphite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2 x 10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7 x 10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7 x 10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0 x 10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4 x 10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4 x 10^{-13}
Hydrogen sulphide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2 x 10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8 x 10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Name: ____

Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	low solubility
	all others	soluble
sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulphide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulphite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

Table of Standard Reduction Potentials with Values

Oxidized Species	↔	Reduced Species	E°/V
$Li^+_{(aq)} + e^-$	↔	$Li_{(s)}$	–3.04
$K^+_{(aq)} + e^-$	↔	$K_{(s)}$	–2.93
$Ca^{2+}_{(aq)} + 2e^-$	↔	$Ca_{(s)}$	–2.87
$Na^+_{(aq)} + e^-$	↔	$Na_{(s)}$	–2.71
$Mg^{2+}_{(aq)} + 2e^-$	↔	$Mg_{(s)}$	–2.37
$Al^{3+}_{(aq)} + 3e^-$	↔	$Al_{(s)}$	–1.66
$Mn^{2+}_{(aq)} + 2e^-$	↔	$Mn_{(s)}$	–1.19
$H_2O_{(l)} + e^-$	↔	$\frac{1}{2}H_{2(g)} + OH^-_{(aq)}$	–0.83
$Zn^{2+}_{(aq)} + 2e^-$	↔	$Zn_{(s)}$	–0.76
$Fe^{2+}_{(aq)} + 2e^-$	↔	$Fe_{(s)}$	–0.45
$Ni^{2+}_{(aq)} + 2e^-$	↔	$Ni_{(s)}$	–0.26
$Sn^{2+}_{(aq)} + 2e^-$	↔	$Sn_{(s)}$	–0.14
$Pb^{2+}_{(aq)} + 2e^-$	↔	$Pb_{(s)}$	–0.13
$H^+_{(aq)} + e^-$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}_{(aq)} + e^-$	↔	$Cu^+_{(aq)}$	+0.15
$SO_4{}^{2-}{}_{(aq)} + 4\,H^+_{(aq)} + 2e^-$	↔	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}_{(aq)} + 2e^-$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^-$		$2OH^-_{(aq)}$	+0.40
$Cu^+_{(aq)} + e^-$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^-$	↔	$I^-_{(aq)}$	+0.54
$Fe^{3+}_{(aq)} + e^-$	↔	$Fe^{2+}_{(aq)}$	+0.77
$Ag^+_{(aq)} + e^-$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^-$	↔	$Br^-_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^-$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7{}^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e^-$	↔	$2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^-$	↔	$Cl^-_{(aq)}$	+1.36
$MnO_4{}^-{}_{(aq)} + 8H^+_{(aq)} + 5e^-$	↔	$Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^-$	↔	$F^-_{(aq)}$	+2.87

Name: ____

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

Names, Formulas, and Charges of Common Ions *(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^-	sulphate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulphate	HSO_4^-
fluoride	F^-	sulphide	S^{2-}
hydroxide	OH^-	hydrogen sulphide	HS^-
hypochlorite	ClO^-	sulphite	SO_3^{2-}
iodide	I^-	hydrogen sulphite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Name: ____

Common Ions

Cations (Positive Ions)

1^{+} charge		2^{+} charge		3^{+} charge	
$NH4^{+}$	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	$Pb2^{2+}$	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

Common Ions *(continued)*

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- $(C_2H_3O_2^-)$	Acetate (or ethanoate)	HS^-	Hydrogen sulphide	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-}	Sulphate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulphide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulphite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulphate
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 sulphate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Name: ____

Final Practice Examination Answer Sheet

Part A: Multiple Choice *(23 marks)*

For each multiple-choice question, shade in the circle that corresponds to your answer. DO NOT circle your answers directly on the examination.

Example: Ⓐ Ⓑ ● Ⓓ

1. Ⓐ Ⓑ ● Ⓓ
2. ● Ⓑ Ⓒ Ⓓ
3. Ⓐ ● Ⓒ Ⓓ
4. Ⓐ ● Ⓒ Ⓓ
5. Ⓐ ● Ⓒ Ⓓ
6. ● Ⓑ Ⓒ Ⓓ
7. Ⓐ Ⓑ Ⓒ ●
8. Ⓐ Ⓑ Ⓒ ●
9. Ⓐ Ⓑ Ⓒ ●
10. ● Ⓑ Ⓒ Ⓓ
11. Ⓐ Ⓑ ● Ⓓ
12. ● Ⓑ Ⓒ Ⓓ
13. Ⓐ Ⓑ Ⓒ ●
14. Ⓐ Ⓑ ● Ⓓ
15. Ⓐ Ⓑ ● Ⓓ
16. ● Ⓑ Ⓒ Ⓓ
17. Ⓐ ● Ⓒ Ⓓ
18. Ⓐ Ⓑ Ⓒ ●
19. Ⓐ ● Ⓒ Ⓓ
20. ● Ⓑ Ⓒ Ⓓ
21. Ⓐ Ⓑ Ⓒ ●
22. Ⓐ ● Ⓒ Ⓓ
23. Ⓐ Ⓑ ● Ⓓ

Part B: Fill-in-the-Blank *(20 marks)*

For each fill-in-the-blank question, write your answer in the space provided that corresponds to the question. DO NOT write your answers directly on the examination.

Equilibrium ***(6 marks)***

1. constant
2. dynamic
3. temperature
4. closed

   system
5. increases

continued

Acids and Bases ***(8 marks)***

6. Arrhenius
7. hydronium
8. acid
9. equal
10. K_{eq}
11. electrolyte
12. hydronium
13. basic

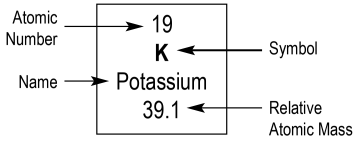
Electrochemistry ***(6 marks)***

14. energy
15. reducing
16. cathode
17. Electrons
18. reductions
19. Electroplating

GRADE 12 CHEMISTRY (40S)

Appendices

Appendix A: Periodic Table of Elements



	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 B: 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)	Irdium	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)	Sulphur	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
Einstienium	(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 C: TABLE OF ELECTRONEGATIVITIES

	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 D: Relative Strengths of Acids Table

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulphuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4×10^{-2}
Sulphurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7×10^{-2}
Hydrogen sulphate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3×10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1×10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0×10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3×10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7×10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1×10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7×10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5×10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6×10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4×10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4×10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4×10^{-7}
Hydrogen sulphide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3×10^{-8}
Hydrogen sulphite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2×10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7×10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7×10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0×10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4×10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4×10^{-13}
Hydrogen sulphide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2×10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8×10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Appendix E: Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg^{2+}, Cu^+, Tl^+, Hg^+	low solubility
	all others	soluble
sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulphide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulphite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

APPENDIX F: TABLE OF STANDARD REDUCTION POTENTIALS WITH VALUES

Oxidized Species	↔	Reduced Species	E°/V
$Li^{+}_{(aq)} + e^{-}$	↔	$Li_{(s)}$	−3.04
$K^{+}_{(aq)} + e^{-}$	↔	$K_{(s)}$	−2.93
$Ca^{2+}_{(aq)} + 2e^{-}$	↔	$Ca_{(s)}$	−2.87
$Na^{+}_{(aq)} + e^{-}$	↔	$Na_{(s)}$	−2.71
$Mg^{2+}_{(aq)} + 2e^{-}$	↔	$Mg_{(s)}$	−2.37
$Al^{3+}_{(aq)} + 3e^{-}$	↔	$Al_{(s)}$	−1.66
$Mn^{2+}_{(aq)} + 2e^{-}$	↔	$Mn_{(s)}$	−1.19
$H_2O_{(l)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)} + OH^{-}_{(aq)}$	−0.83
$Zn^{2+}_{(aq)} + 2e^{-}$	↔	$Zn_{(s)}$	−0.76
$Fe^{2+}_{(aq)} + 2e^{-}$	↔	$Fe_{(s)}$	−0.45
$Ni^{2+}_{(aq)} + 2e^{-}$	↔	$Ni_{(s)}$	−0.26
$Sn^{2+}_{(aq)} + 2e^{-}$	↔	$Sn_{(s)}$	−0.14
$Pb^{2+}_{(aq)} + 2e^{-}$	↔	$Pb_{(s)}$	−0.13
$H^{+}_{(aq)} + e^{-}$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}_{(aq)} + e^{-}$	↔	$Cu^{+}_{(aq)}$	+0.15
$SO_4{}^{2-}{}_{(aq)} + 4\ H^{+}_{(aq)} + 2e^{-}$	↔	$H_2SO_{3(aq)} + H_2O_{(l)}$	+0.17
$Cu^{2+}_{(aq)} + 2e^{-}$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^{-}$		$2OH^{-}_{(aq)}$	+0.40
$Cu^{+}_{(aq)} + e^{-}$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^{-}$	↔	$I^{-}_{(aq)}$	+0.54
$Fe^{3+}_{(aq)} + e^{-}$	↔	$Fe^{2+}_{(aq)}$	+0.77
$Ag^{+}_{(aq)} + e^{-}$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^{-}$	↔	$Br^{-}_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^{+}_{(aq)} + 2e^{-}$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7{}^{2-}{}_{(aq)} + 14H^{+}_{(aq)} + 6e^{-}$	↔	$2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^{-}$	↔	$Cl^{-}_{(aq)}$	+1.36
$MnO_4{}^{-}{}_{(aq)} + 8H^{+}_{(aq)} + 5e^{-}$	↔	$Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^{-}$	↔	$F^{-}_{(aq)}$	+2.87

APPENDIX G: 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^-	sulphate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulphate	HSO_4^-
fluoride	F^-	sulphide	S^{2-}
hydroxide	OH^-	hydrogen sulphide	HS^-
hypochlorite	ClO^-	sulphite	SO_3^{2-}
iodide	I^-	hydrogen sulphite	HSO_3^-
iodate	IO_3^-	thiocyanate	SCN^-

Appendix H: Common Ions

Cations (Positive Ions)

1^+ charge		2^+ charge		3^+ charge	
$NH4^+$	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	$Pb2^{2+}$	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 sulphide	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-}	Sulphate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulphide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulphite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulphate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocynate	**3^- charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulphate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

APPENDIX I: WORKING WITH SIGNIFICANT FIGURES

Introduction

If our calculator calculations are done 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 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 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 metre 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 an estimation 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.

Zeros (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 mathematics course.

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

   Examples:

   374 (three significant digits)

   8.1 (two significant digits)

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

   Examples:

   50407 (five significant digits)

   8.001 (four significant digits)

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

   Examples:

   0.54 (two significant digits)

   0.0098 (two significant digits)

4. *Trailing zeros 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. The truth is that they are not the same thing. 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 zeros are not significant unless stated so.*

   Examples:

   37.000 (5 significant digits, because the decimal point identifies that the trailing zeros are significant)

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

   To show when trailing zeros are significant, we use **scientific notation**. All zeros 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 number 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 zeros.

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

2.5×3.42

The answer to this problem would be 8.6, rounded from the calculator reading of 8.55. Why? Because 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

3.10×4.520?

You may have said two. This is too few. A common error is to look at a number like 3.10 and think that 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 least 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 as follows:

If a decimal point is *present*, ignore zeros on the ***P**acific* (left) side of a number. If the decimal point is *absent*, ignore zeros on the ***A**tlantic* (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.

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, maintaining 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 × 6.085 × 2.1
   d. 8432 ÷ 12.5

Answers:

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, maintaining the correct number of significant figures in your answers.
   a. 23.1 + 4.77 + 125.39 + 3.581

      To one decimal place, the answer is 156.8 (three significant figures).

   b. 22.101 – 0.9307

      To three decimal places, the answer is 21.170 (five significant figures).

   c. 2.33 × 6.085 × 2.1

      To two significant figures, the answer is 30.

   d. 8432 ÷ 12.5

      To three significant figures, the answer is 674.

Significant Figures 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.

Source: Manitoba Education. *Grade 11 Chemistry: A Course for Independent Study: Field Validation Version*. Winnipeg. MB. 2010. Module 2: 29-36.

Notes

GRADE 12 CHEMISTRY (40S)

Bibliography

BIBLIOGRAPHY

Buthelezi, Thandi, et al. *Chemistry: Matter and Change: Student Text*. New York, USA: The McGraw-Hill Companies, Inc., Glencoe Science, 2008.

Buthelezi, Thandi, et al. *Chemistry: Matter and Change: Teacher Classroom Resources*. New York, USA: The McGraw-Hill Companies Inc., Glencoe Science, 2008.

Chastko, Audrey, et al. *Inquiry into Chemistry*. Toronto, ON: McGraw-Hill Ryerson, 2007.

LeMay, H. Eugene, Jr., et al. *Chemistry: Connections to Our Changing World*. 2nd ed. Upper Saddle River, NJ: Prentice-Hall Inc., 2000.

Manitoba Education. *Chemistry – Grade 12: A Foundation for Implementation Draft / Unedited Version*. Winnipeg, MB: Manitoba Education, 2007.

North Carolina School of Science and Mathematics (NCSSM). *Chemistry Files: Animations and Videos* (CD). North Carolina, U.S.A.: North Carolina School of Science and Mathematics, 2013.

North Carolina School of Science and Mathematics (NCSSM). *Teachers' International Graphics Educational Resource (T.I.G.E.R.)*. North Carolina, U.S.A.: North Carolina School of Science and Mathematics, 2013.

Phillips, John S., et al. *Chemistry: Concepts and Applications*. New York, USA: The McGraw-Hill Companies Inc., Glencoe/McGraw-Hill, 2000.

Spence, D. "Principal Energy Levels and Sub-orbitals." *Mr. Spence's Science Website*. (21 Feb. 2013). http://sun.menloschool.org/~dspence/chemistry/atomic/config2.html

Wilbraham, Antony, et al. *Chemistry*. Needham, MA: Prentice-Hall Inc., 2002.

van Kessel, Hans, et al. *Chemistry 12*. Toronto, ON: Thomson Nelson, 2003.

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