CHEMISTRY

TEACHER'S GUIDE

GRADE 12

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FEDERAL DEMOCRATIC REPUBLIC OF ETHIOPIA
MINISTRY OF EDUCATION

Published E.C. 2002 by the Federal Democratic Republic of Ethiopia, Ministry of Education, under the General Education Quality Improvement Project (GEQIP) supported by IDA Credit No. 4535-ET, the Fast Track Initiative Catalytic Fund and the Governments of Finland, Italy, Netherlands and the United Kingdom.

The Ministry of Education wishes to thank the many individuals, groups and other bodies involved – directly and indirectly – in publishing this textbook and the accompanying teacher guide.

Developed and Printed by

STAR EDUCATIONAL BOOKS DISTRIBUTORS Pvt. Ltd.
24/4800, Bharat Ram Road, Daryaganj,
New Delhi – 110002, INDIA
and
ASTER NEGA PUBLISHING ENTERPRISE
P.O. Box 21073
ADDIS ABABA, ETHIOPIA

Under GEQIP Contract No. ET-MoE/GEQIP/IDA/ICB/G01/09.

978-99944-2-041-4

INTRODUCTION

According to the Educational and Training Policy of the Federal Democratic Republic of Ethiopia, the second cycle of the secondary education and training will enable your students to choose subjects or areas of training which will prepare them adequately for higher education and for the world of work. The study of chemistry at this second cycle, Grades 11 and 12, prepare your students academically as well as psychologically for the nation building. Studying chemistry provides students not only with specific concepts and theories in chemistry, but also with tools, confidence and attitudes for constructing their future prosperous society. Besides learning to think effectively and efficiently, your students come to understand how chemistry deals with the daily and routine lives of theirs and the people at large. Your students are also expected to realize the changing power of chemistry and its significance nationally as well as internationally.

To materialize the above stated major goals, encourage your students to apply high-level reasoning, and values to their daily life and also to their understanding of the social, economic, and cultural realities of the surrounding context. In turn, this will help the students to actively and effectively participate in the wider scope of the development activities of their nation.

At this cycle, the students are highly expected to gain solid knowledge of the fundamental theories, rules and procedures of chemistry. It is also expected that they should develop reliable skills for using this knowledge to solve problems independently and in groups. To this end, the specific objectives of chemistry learning at this cycle are to enable the students to:

- ✓ gain a solid knowledge of chemistry.
- ✓ appreciate the changing power, dynamism, structure and elegance of chemistry.
- ✓ Apply chemistry in their daily life.
- ✓ understand the essential contributions of chemistry to the fields of engineering, science, agriculture and economics at large.

Recent research gives strong arguments for changing the way in which chemistry has been taught. The traditional teaching-learning paradigm has been replaced by active, participatory and student-centered model. A student-centered classroom atmosphere and approach stimulates student's inquiry. Your role as a teacher in such student-oriented approach would be a mentor who guides the students construct their own knowledge and skills. A primary goal when you teach a fundamental theory is for them to discover the concept by themselves, particularly as you recognize threads and patterns in the data and theories that they encounter under the teacher's guidance and supervision.

You are also encouraged to motivate your students to develop personal qualities that will help them in real life. For example, encourage students' self confidence and their confidence in their knowledge, skills and general abilities. Motivate your students to express their ideas and observations with courage and confidence. As the students develop personal confidence and feel comfortable on the subject, they would motivate addressing their material to groups and to express themselves and their ideas with strong conviction. Support students and give them chance to stand before the class and present their opinion, observation and work. Similarly, help the students by creating favorable conditions for students to come together in groups and exchange views and ideas about what they have investigated and about the material they have read. In this process, the students are given opportunities to openly discuss the knowledge they have acquired and to talk about issues raised in the course of the discussion.

Teamwork is one of the acceptable ways of approach in a student-centered classroom setting. For example, some experiments are performed by more than one student. Each student has a crucial role – one student might be responsible for carefully handling and mixing chemicals and another student may make quick and accurate measurements during the process.

This teacher's guide helps you only as a guide. It is very helpful for budgeting and breaking down your teaching time as you plan to approach specific topics. The guide also contains procedures to manage class activities, group discussions and reflections. Answers to the review questions are indicated at the end of each topic.

Every section of your teacher's guide includes student-assessment guidelines. Use them to evaluate your students' work. Based on your class's reality, you will give special attention to students who are working either above or below the standard level of achievement. Do an active follow up for each student's performance against the learning competencies presented in the guide. Be sure to consider both the standard competencies and the minimum competencies. ***Minimum requirement level*** is not the ***standard level of achievement***. To achieve the standard level, your students must fulfill all of their grade-level's competencies successfully.

When you identify students who are working either below the standard level or the minimum level, arrange them extra support. For example, you can give them supplementary presentations and reviews of the materials in the class. Extra time to study and activities to those who are performing below the minimum level is commendable. You can also encourage high-level students with advanced activities and extra exercises.

Some helpful references are listed at the end of this teacher's guide. For example, if you get an access for internet it could be a rich resource for you. Searching for web sites is well worth your time as you browse on the subject matter you need. Use one of the many search engines that exist – for example, Yahoo and Google are widely accepted.

Do not forget that, although this guide provides many ideas and guidelines, you are encouraged to be innovative and creative in the ways you put your students into practice. Use your own full capacity, knowledge and insights in the same way as you encourage your students to use theirs.

GENERAL INFORMATION TO THE TEACHER

The students' text is designed and prepared based on the participatory approach of the teaching – learning process. At present, it is believed that students should gain most of their knowledge from the teaching – learning process on their own and some from the teacher. The teacher is expected to give guidance and the necessary assistance, play a role as facilitator, harmonize concepts, provide students with materials required, create a conducive atmosphere for the teaching – learning process and evaluate of students' performance. The teacher needs to assist students to discover facts, realize concepts, develop skills in performing experiments, solving problems etc. So, he/she should not dominate the teaching – learning process by giving lecture or explaining concepts throughout the period.

Thus, whenever you have contact with your students, you need to plan how to promote active – learning. The following information will help you understand what you are expected to do before and during the entire teaching – learning process.

1. Organizing groups

You need to organize different groups in each section you are going to teach during your first contact with the students. To do so, you better have the list of all students in each section. You may organize the groups based on their seats, or on their ability as slow learners, medium and fast learners or by mixing them. After organizing groups, give them group numbers as group 1, group 2 etc. and register the names of students in each group. Every group needs to have a group leader and a secretary to jot down the main points during discussions. The groups as well as their members need not be permanent throughout the year. You can reorganize groups whenever necessary. You can do so per semester or mid-semester or even per month or two months etc.

2. Discussion

In all units, sections and subtopics, there are activities suggested for students to help them discuss and discover concepts. When you allow them to discuss points in each activity:

- follow up how every student participates in the discussion.
- be part of the discussion in some groups for a few minutes and see how the discussion among students is going on.
- give assistance and guidance when students are in need.

  - give them hints when they face difficulties or have questions on the points suggested in the activities.
  - ask questions related to the points in the activity to facilitate the interaction among students during discussion.

3. Presentation

Students are expected to present:

a. the concepts they gained during discussion in each activity in all units.
b. their observation and analysis after performing experiments in groups to the class.
c. the content prepared a specific topic. So you need to give emphasis to the following points in order to maximize student participation.
    a. Groups should present their opinion turn by turn. For example, if you allow group 1 and group 2 to make a presentation on activity 1.1, the following groups 3 and 4 or others will present activity 1.2 etc.
    b. Whenever a group gets the chance to make presentation for the second or third time, let other members of the group accomplish the task. Do not allow the same student from the same group to do so.
    c. Give the opportunity to the rest of the class to ask questions or give their comments on the presentation of a particular group.

4. Experiment

Several experiments are suggested in the first three units. Most of these experiments should be performed by students. So, you are expected to accomplish the following tasks before or when students carry out the experiment.

a. To carry out the experiment by yourself before allowing students to do it.
b. To prepare chemicals and apparatus required for the experiment.
c. To give instruction on how students should handle chemicals and apparatus during every experiment.
d. To provide materials they need for the experiment.
e. Assist them whenever they have questions or difficulties in understanding the procedures suggest0ed for the experiment.
f. Give instructions that students should perform the experiment only based on the procedures suggested for it.
g. Never allow them to conduct an experiment on their own other than the one they are supposed to do during the period.

h. Make them write a laboratory report in groups, present their observation to the rest of the class or submit it to you for correction as suggested in the students' text.

i. Make sure that every student in each group participates in the experiment.

5. Harmonizing Concepts

You are not expected to lecture throughout the period on most of the contents in the students' text. Your major role is harmonizing concepts suggested by students during presentations after discussing activities or performing an experiment with those they are expected to know. So, you need only to build a mini – lecture.

The concepts intended for students to discover in all activities, and answers to questions on the observation and analysis part of all experiments, are included as short notes in the subject matter presentation part of every section in this teachers' guide. So you are advised to use them. While harmonizing concepts in a mini – lecture, you better include other contents of the topic that have not been covered when students discuss activities.

6. Continuous Assessment

Previously, the performance of a student has been assessed in terms of his/her achievements in quizzes, tests, homework, mid – semester and semester final examinations. Although these evaluation techniques are useful tools for the assessment, they may not give a clear picture of the performance of a student. Therefore, a student's work should be assessed throughout every topic, section and unit as well as during each period. So, you need to have a record of every student's work as a student performance list. You can make a record about each student in the performance list, based on the following points.

- Involvement in discussions.
- Participation in presentations after discussion.
- Participation in answering questions during the process of harmonizing concepts or stabilization.
- Role of the student in performing experiments.
- Role of the student in presenting concepts gained from the experiment.
- Presentation of the project work.
- Presentation of research and writing.
- Presentation of topics given to the group as homework.

- Answering questions accordingly given as
  - ✓ class work
  - ✓ homework
  - ✓ quizzes
  - ✓ tests
  - ✓ mid – semester and semester final examinations

Here, it is very important to note that the assessment system is continuous assessment. That is, every performance of the student during the teaching-learning process should be given value and contribute its own share, as do quizzes, tests, mid-semester and semester final examinations, to the semester total. You are empowered to decide the percent of the contribution. However, your decision should not violate either the policy of the Ministry of Education or that of the Education Bureau of the regional state or that of your school.

7. Additional Questions

Some questions are given in this teachers' guide in each section before the answers to the exercises in the section. Use the questions indicated by an asterisk (*) for students working below the minimum requirement level, while students working above the minimum requirement level can attempt all of them. Give these questions as class work for fast learners after they complete their work during each period so that they will not sit idle and the period will not be boring for them.

8. Giving Note

You are not expected to write notes on the board related to the contents in each section. You need to give short notes on those contents left for students to discover after discussing the suggested activities and performing experiments. Be sure to offer any note that is available in the teacher's guide, but not in the students' text. However you can write short notes related to the main points as you harmonize concepts. Tell students how they can take notes, either from the text or during the teaching learning process. Tell them the main points they should emphasize, in taking notes from the text. Also tell them to jot down the main points as fast as they can as you harmonize concepts or give a mini-lecture.

9. Answers to Exercise

In all units, the answers to the suggested exercise are given at the end of each section, and answers to the review exercises in each unit at the end of the unit. So you can refer to them whenever you are in need.

10. Suggested methodologies

Teaching all contents of grade 10 chemistry requires implementing active learning methodologies. Active learning involves providing opportunities for students to participate in meaningful talk and to listen, write and reflect on the content, ideas, issues and concerns of an academic subject. It is more of a student activity. The teacher is a facilitator. The teacher guides and directs the students.

Rationale for active learning

- an increase in academic achievements
- an increase in critical thinking skills
- increased student retention
- a more positive attitude toward the subject matter
- improvement in communication skills

There are many methods that can be used to implement active learning. However, all of them are not suitable for teaching chemistry. So, some of the methodologies that can be used to promote active learning in teaching chemistry at this level are suggested as follows.

A. Gapped lectures

You divide your lecture into small sections (lecture for a period of 15 minutes) and give the students a quick activity of 5 to 10 minutes. After the activity, you proceed with another 15 minutes lecture followed by another activity. The activities usually emphasize the concepts included in the lecture. For example, you can apply this methodology to teach concentration of solutions.

B. Cooperative (collaborative) learning

This is a form of group work and it is helpful in group project work and group assignments. This can be applicable for students in doing their group assignments or in doing suggested project work. For example, in types of solutions.

C. Group discussion

Is a simple interaction pattern in which 4 – 6 students work together on a given task and produce a written work or presentation. This method can be used in all sections and units at this level.

D. Presentation

This is an activity where students present a topic in front of their classmates. This can be done individually or as a group. For example, in all units after group discussion.

E. Demonstration

This is a method where the teacher shows the students how something is done. For example, in unit 4 (electrochemistry)

F. Experiments

It usually involves a very specific and controlled method of procedures, where results are usually recorded.

G. Concept map

It is a visual representation of ideas on any given topic. Students write the topic at the center of the page and then divide it into subtopics from which smaller branches will go off in different directions. For example, to teach colligative properties of solutions.

H. Question and answer (inquiry)

When this method is used, the teacher lectures and asks questions periodically relating to the information being given. For example, to teach thermochemistry.

I. Spider diagram

Students write a topic at the middle and write ideas related to the topic around the topic and draw a line connecting each idea to the central idea. For example, to teach ways of expressing concentrations of solutions.

J. Visual-based active learning

This method helps students learn using real object models, pictures, drawings and charts. For example, you can use this method in unit three to teach entropy.

K. Brain storming

This is an activity in which students write everything they know or think about a given topic. The ideas might be right or wrong. This can be done individually, in pairs, small groups or as a whole class with the teacher or a student recording the ideas on the board. This method is used to find out what students already know on a topic before you start teaching. For example this method can be used teach intensive properties.

You can use the following websites to get more information on active-learning methodologies.

1. **http://www.ntlf.com/html/lib/bib/91-9dig.htm**
2. **http:ctl.byu.edu/active-learning-techniques/**
3. **http:pdfcast.org/pdf/strategies-to-incorporate-active-learning-into-online-teaching**
4. **http://ijklo.org/volume5/IJELLOv5p215-232Pundak669.pdf**

11. Motivation of students and its importance

Motivation of students means getting students to exert a high degree of effort in their learning activities. The teacher is expected to motivate the students to create a conducive atmosphere for the teaching learning process. To motivate students, the teacher needs to encourage them to get ready for the lesson, appreciate students for their attempts in answering questions or any other activity they perform during the teaching-learning process and give them recognition. Motivating students helps the teacher.

- to pass information to students according to the plan
- to make students active participants
- make students realize concepts easily
- make his/her teaching interesting
- achieve the desired goals etc.

Motivation also helps students to

- follow the lesson attentively
- increase their participation
- enhance their understanding
- develop interest in the subject
- achieve good results in their performance

Implementing active learning methodologies has a role of its own in motivating teachers as well. It is not as tiresome as that of lecturing although, the teacher has a lot of tasks to accomplish when applying the methods. Using active learning methodologies during the teaching learning process motivate the teacher to:

- enjoy friendly and interesting relationships with students.
- develop new teaching skills by practicing the new teaching techniques, observing their results, and contrasting them with those of the old method of lecture-based teaching.
- become more interested in the teaching profession. For example, it is interesting and satisfying to develop new skills. The teaching-learning approach guides the teacher, helping him or her to develop professionally.
- investigate each student's talents and creativity. In this way, the teacher learns more about the age group of the students he or she teaches. This process is interesting in itself and helps the teacher develop professionally.
- guide students individually as they learn on their own. In this way, the teacher learns more about the dynamics of learning and also of teaching.
- actively engage in furthering the students' development. Because the students develop important social skills and attitudes, as well as increasing their knowledge and learning skills, the teacher has the satisfaction of contributing to their community and therefore to the country as a whole.
- expand his or her own creativity by developing appropriate presentations and assembling the apparatus and the local materials required for demonstrations and experiments.
- develops a greater interest in the teaching profession. As he or she assumes direct responsibility for each student's development.

Table of Contents

Page

UNIT 1 SOLUTION

UNIT OVERVIEW

Total Periods allotted: 30 periods

This unit deals with solutions. Since solutions are homogenous mixtures, the unit starts with types of mixtures. Section 1.1 deals with two types of mixtures: homogeneous and heterogeneous mixtures. The types of solutions are described in section 1.2 as gaseous solutions, liquid solutions and solid solutions. Section1.3 deals with solution processes in terms of interparticle forces of interaction. Illustrative examples and activities are given for liquid-liquid, gas-liquid and solid-liquid solutions. The effect of temperature and pressure on solubility is discussed in section 1.4. Ways of expressing concentration of solutions are explained in section 1.5. This includes percent by mass, mole fraction, molarity, molality and molality and parts per billion. Preparations of solution are discussed in section 1.6. Section 1.7 describes solution stoichiometry. Reactions in solutions are described in terms of molecular and ionic equations in section 1.8.

Section 1.9 examines some physical properties of solutions called colligative properties, and phenomena related to them. In each section of the unit there are suggested activities, projects and experiments. This is aimed at helping both students and teachers to practice the active-learning approach. The methodologies suggested to teach the contents in this unit are gapped lecture, group discussion, question-and-answer, experiment and others mentioned in the sub-topics in each unit.

Unit Outcomes

After completing this unit, students will be able to:

- *know the types of solutions;*
- *understand the solution formation process, the rate of solution, the heat of solution and solubility;*
- *describe the dependence of solubility on temperature and pressure of solution;*
- *know how to solve problems involving concentration of solution and to express the result in various units;*
- *describe the behavior of ionic solute in solutions that are unsaturated, saturated and supersaturated, using the concept of equilibrium;*

- *prepare solutions of required concentrations by dissolving a solute or diluting a concentrated solution;*
- *understand the relationship between the colligative properties of solutions and know how to solve problems involving colligative properties of solutions;*
- *describe scientific inquiry skills used in this unit: observing, classifying, comparing and contrasting, measuring, asking questions, drawing conclusions, applying concepts and problem solving.*

Main Contents

1.1 Homogeneous and Heterogeneous Mixtures
1.2 Types of Solutions
1.3 The Solution Process
1.4 Solubility as an Equilibrium Process
1.5 Ways of Expressing Concentrations of Solution
1.6 Preparation of Solutions
1.7 Solution Stoichiometry
1.8 Describing Reactions in Solution
1.9 Colligative Properties of Solutions

1.1 HOMOGENEOUS AND HETEROGENEOUS MIXTURES

Period Allotted: 1 period

Competences

At the end of this section, students will be able to:

- *define the terms: mixture, homogeneous and heterogeneous mixtures, solute, solvent, solution;*
- *distinguish between homogeneous and heterogeneous mixtures;*
- *describe suspensions and colloids.*

Forward Planning

Read the text book and other necessary references to familiarize yourself with the types of mixtures. Make sure that the necessary materials are available to do experiments 1.1 and 1.2. Make a plan for managing students when they perform experiment in groups.

Teaching Aids

Beakers, test tubes, watch glasses, stirrers, table salt, water, iron filings, sulphur powder, oil, alcohol, three flasks, sand, gentian violet and bile.

Subject Matter Presentation

Use gapped lectures and experiment as the methodologies for this section.

Start introducing this section by asking the students to state the two major types of mixtures. After they respond, help the students to do the experimental activity. Let them do Experiment 1.1 and differentiate homogenous and heterogeneous mixtures.

Based on the experiment, let them suggest the definitions of the two types of mixtures. Make sure that they classified a solution of table salt and water, alcohol and water as homogenous mixtures and combination of iron filings and sulphur powder, oil and water as heterogeneous mixtures. Then, tell them that in this unit, only homogeneous mixtures will be studied, because a solution is a homogeneous mixture. Ask your students this question and discuss it: "Why are solutions necessary after all?" Here you can mention the reasons why we study solutions. For example, the occurrences of many biological reactions in living things exist in solutions and most chemical reactions are seen in solution. Solutions usually have higher boiling points and lower freezing points than pure solvents. The difference between suspensions, colloids and solutions should be clear to the students.

Let students perform Experiment 1.2 under your supervision. Ask the students to dissolve sand, gentian violet and bile in water in three different flasks. Set them aside and see if they become clearer on standing for some time. Tell them that they need to decant after 30 minutes. Ask them whether they can see changes after further standing.

Then let them pass light through each mixture and see which one scatters light without separating on standing. Then ask what they have observed. This will help them distinguish the differences between solutions, suspensions and colloids.

After completing the experiment, let them write a laboratory report in groups and one or two groups present their observations to the class. Check whether or not each report is in accordance with the following facts of the experiment.

a. The contents in beaker A separate in two layers on standing.

b. The contents in beaker C scatter light without separating

c. Beaker B contains a solution, while beaker C contains a colloid.

In case of (d) tell them the differences between suspension, solution and colloid based on their presentation to avoid confusion or misunderstanding of concepts.

Emphasize the components of solutions: solute and solvent. Briefly explain solute and solvent by using illustrative examples such as dissolving sugar in water. Ask the students which of these materials is a solute and which is a solvent.

Assessment

Assess each student's work throughout section 1.1. Carefully check how every student is working in the teaching-learning process. In a permanent performance list, record each student's participation.

- in performing experiments 1.1 and 1.2
- in presenting the observation and analysis of the experiments
- in answering questions raised during the process of harmonizing concepts.

Give them exercise 1.1, check their work and record their achievements. From your records, see whether or not the competencies suggested for section 1.1 are achieved by most of the students.

Praise students working above the minimum requirement level and assist those working below the minimum requirement level. To assess the attainment of the minimum requirement level, ask the students a variety of questions. For example, you can ask them about the differences between homogeneous and heterogeneous mixtures, colloids and solutions. You can also ask them to define the terms solute, solvent and solution.

Additional Questions

1. Are colloids solutions or mixtures?

* 2. Classify the following as colloids, solutions or suspension

a.	Fog	b.	smoke
c.	paint dissolved in water	d.	Cell fluid
e.	firefighting foams	f.	$Al(OH)_3$ used as an antiacid.

Answers to Additional Questions

1. Mixtures
2. All are colloids except 'f' which is a suspension.

Answers to Exercises

Exercise 1.1

a. A homogenous mixture is a combination of two or more substances that has the same composition throughout and has no visible boundary.

b. A heterogeneous mixture is a combination of two or more substances that has no uniform composition throughout and contains one or more visible boundaries between the components.

c. A solute is a component of a solution present in a smaller amount than the solvent.

d. A solvent is either a components of solution that is present in a large amount or the component that determines the physical state of the solution.

e. A colloid is a mixture containing particles that are small enough to not settle down ward.

f. A suspension is a heterogeneous mixture containing particles that are visible either to the naked eye or with the help of a magnifying glass and that can settle down wards when the mixture is allowed to stand.

1.2 TYPES OF SOLUTIONS

Periods Allotted: 2 periods

Competencies

At the end of this section, students will be able to:

- *explain different types of solutions;*
- *give examples of each type of solution;*
- *present to the class a report on how jewellery gold is made after a visit to a nearly goldsmith shop .*

Forward Planning

Prepare yourself by reviewing the three states of matter, solids, liquids and gases, so that you can connect this idea with types of solutions. Plan how to manage students during discussions and the presentations after discussion. Find out whether or not there is a goldsmith near your school so that you may arrange a visit. Read the teacher's guide on this section to get more information about Activity 1.1.

Subject Matter Presentation

To teach the contents in this section use group discussion, collaborative learning and gapped lecture as methodologies.

Let the students list some of the solutions that they encounter in their daily lives which are gaseous solutions, liquid solutions and solid solutions.

Form groups of four students and start teaching with Activity 1.1. This activity is aimed at helping the students to know the types of solutions. Let the students discuss Activity 1.1 for a few minutes and present it to the class. Make sure that they include gaseous solutions, liquid solutions and solid solutions in their presentations. Let them construct a table and fill it with necessary information as follows:

State of solute	**State of solvent**	**State of solution**	**Type of solution**	**Examples**
Gas	Gas	Gas	Gas-gas	Air
Gas	Liquid	Liquid	Gas-liquid	Oxygen in water
Liquid	Liquid	Liquid	Liquid-liquid	Alcohol in water
Solid	Liquid	Liquid	Solid-liquid	Salt in water
Solid	Solid	Solid	Solid-solid	Bronze (Cu and Sn)

Since students may not able to give examples of solid solutions as readily as they did for the other types of solution, use alloys as examples. Ask the students if they know how jewellery gold is made. It is an alloy of cheaper metals dissolved in gold, after both solute and solvent are melted. The homogeneous mixture (solution) then solidifies upon cooling.

In the case of **Project 1.1**, you could arrange a visit to a nearby goldsmith for a demonstration and ask the students to write a short report about what they have observed. Let them discuss their observations when they are back in class. Harmonize their discussion.

Assessment

Assess each student's work throughout the sub-unit (types of solution). You can do this by closely watching how every student is doing in this section and recording their performances. You can make records based on the students participation

- in discussing Activity 1.1
- in presenting opinions after discussion
- in performing the project work and presenting findings from the project

Give them exercise 1.2 as class work or homework. Check their work and record their performances. In addition to that, you may ask the students to give examples of solid-liquid, gas-liquid, liquid-liquid and solid-solid solutions. In connection with this, you may raise the question of how carbon dioxide is used to prepare alcoholic beverages such as beer and nonalcoholic beverages such as ambo mineral water. See how the students answer these questions.

From your records and observation, check whether or not the competencies suggested for section 1.2 are achieved. Praise students working above the minimum requirement level. Arrange extra lesson time or give them additional exercise to those working below the minimum requirement level

Additional Questions

* 1. What type of solution is steel?

2. Dental amalgam contains mercury and silver. To which class of solutions does it belong?

Answers to Additional Questions

1. solid-solid.
2. liquid-solid.

Exercise 1.2

1. a. air, natural gas
   b. ambo water, Coca-Cola, beer
   c. alcohol in water, benzene in carbon tetracholoride
   d. sugar in water, copper sulphate in water
   e. brass (Zn/Cu), bronze (cu/Sn)
2. a. gas + gas
   b. gas + liquid or liquid + liquid or solid + liquid
   c. solid + solid or liquid + solid (in the case of alloys of mercury with other metals)

1.3 THE SOLUTION PROCESS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *explain how the "like dissolves like" rule depends on interparticle forces of interactions and predict relative solubility;*
- *define rate of dissolution;*
- *define heat of solution, solvation energy and hydration energy;*
- *apply the concept of heat of solution to the solution of ammonium nitrate and sodium hydroxide crystals;*
- *explain how heat of solution is influenced by the interparticle interaction forces.*

Forward Planning

Read about the types of solution from the student text and other reference materials as well as from this guide thoroughly and make appropriate preparation. Design a plan that shows which contents, activities and experiments you are going to deal with during each period so that the whole content in this section can be covered within five periods. Indicate the time duration you allot for every activity you perform in each period. You need to plan how to manage students during discussion, presentation and experiments. Make sure that the diagrams you use to teach this section and also apparatus and chemicals required to perform Experiment 1.3 are available in the school laboratory. Carry out the experiment yourself before you allow students to do it.

Teaching Aids

- Refer to the students text for the materials required to perform Experiment 1.3.
- Flip charts of Figure 1.3 and Figure 1.5.

Subject Matter Presentation

Liquid solutions and interparticle forces of attractions

Use group discussion, gapped lecture and visual–based active learning as methodologies for this topic.

You may start the lesson by using Activity 1.2. Let the students discuss the activity for a few minutes in groups and then have students from two different groups suggest their opinions. This helps the students to understand why some substances do not mix while some others are miscible with one another. In the activity, students can observe that oil and water, and also benzene and water are not miscible because oil and benzene are nonpolar while water is polar. But water and ethanol are miscible because they are both polar.

After harmonizing concepts from Activity 1.2, write the following questions on the blackboard.

1. What are the three types of interparticle forces that determine the tendency of solute particles mix with particles of the solvent?
2. Which interpartice forces should be stronger to make it easier to dissolve a solute in a solvent?
3. What do you understand from the principle “like dissolves like”?
4. Benzene is soluble in carbontetrachloride but not in water. Why?

Let the students discuss these questions in their groups for a few minutes and then have two students from different groups suggest their opinions.

Then, continue by presenting your mini-lecture on *Liquid Solutions and Inter-particle Forces of Attraction.* In your lecture, it is good to use Figure 1.3 drawn on a flip chart. Explain that the fundamental process leading to the dissolution of substances is the interaction between solute and solvent molecules (intermolecular interaction). The key point in this section is that, in order for substances to be appreciably soluble in each other, they must possess similar intermolecular attractive forces. Substances that attract each other very strongly tend to congregate and separate from those to which they are weakly attracted.

The intermolecular forces are identified as solute-solute, solvent-solvent, and solute-solvent. These interactions vary greatly with the physical state of the solute and the solvent. For this reason, solutions of liquids in liquids and solids in liquids are separately considered in the text book.

Solutions of solids in liquids

It is advisable to use group discussion, gapped lectures and visual–based active for this topic.

To look at the solution of solids in liquids, let the students do Activity 1.3, which enables them to identify whether or not common salt (NaCl) and common sugar are soluble in both water and oil. Let them discuss their observations. Let students present their generalizations to the class.

Based on their presentations, harmonize the students' generalizations with the basic facts in the lesson to be taught. In your lecture, it is good to use Figure 1.5 drawn on a flip chart.

Let your students remember that, when a solute particle is placed in solution, it becomes solvated. In other words it becomes surrounded by solvent molecules, to which it is attracted. The case of soaps and detergents could be used as a practical example to explain the dissolution of solids in liquids. Soaps and detergents work on the principle of "like dissolves like". Nonpolar tails of soap anions dissolve in oil and grease globules. The polar heads keep the particles suspended in water so they can wash away.

As you proceed with your lesson delivery on *The Rate of Dissolution*, you may involve your students by letting them read the note in the text and by having them present their findings to the class.

Energy change in the solution process

Use experiment and question and answer methods to teach this topic.

Let your students notice that the formation of solution is often accompanied by the absorption or release of heat. The energy absorbed or liberated during the formation of a solution is called the heat of solution ($\Delta H_{solution}$). Experiment 1.3 demonstrates the heat of solution of NaOH and $NH_4 NO_3$. Help the students to do this experiment and to write a report of the experiment. (You may arrange a separate laboratory class for this experiment).

At this point, try to connect the nature of the heat of solution to the type of inter-molecular interaction. For liquid solutions formed from a solvent A and solute B, an ideal solution results when the A – B interactions are the same as the A – A and B – B interactions. For this solution, the heat of solution ($\Delta H_{solution}$) is zero. When the A – B interactions are greater than the A – A and B – B interactions, particles in solution are more strongly attracted to each other than when they are in pure. The two components mix to form a solution, and the extra energy is given off as heat, ($\Delta H < 0$). This solution process is **exothermic**. Test tubes containing solutions of NH_4Cl and KNO_3 are cold, so the solution process is endothermic. In contrast, the test tube containing Na_2CO_3 is warm, so the solution process is exothermic.

On the contrary, when the A – B interactions are weaker than the A – A and B – B interactions, $\Delta H_{solution} > 0$ and the solution process is **endothermic**.

Hydration of Ionic Solids in Water

Use group discussion and gapped lecture as methodologies to teach this topic.

After introducing the topic for the lesson, tell the students that the solubility of an ionic solid depends on the energy of hydration of ions and the lattice energy. Explain the relationship between the magnitude of lattice energy and the solubility of ionic solids. Help them to recall the factors influencing the magnitude of lattice energy.

Next, proceed to Activity 1.4. This activity enables students realize that solution formation is either exothermic or endothermic. Let the students do Activity 1.4. On dissolving each compound in three different test tubes, the one which teals hot on touching the test tube is exothermic and the one which is cold is endothermic. Because the test tubes containing solutions of NH_4Cl and KNO_3 are cold, the solution process is endothermic. The test tube containing Na_2CO_3 is warm, so the solution process is exothermic.

Activity 1.5 illustrates the relationship between solubility and lattice energy. Have them discuss this activity for a few minutes and some groups present their opinion. After the presentation, harmonize concepts asking the following information.

Lattice energies depend on the charge on the ions and also on the distance between the centers of the neighbouring positive and negative ions. As the magnitude of the charge on the ions increases, the lattice energy also increases. For this reason, you can expect substances with single charged ions to be more comparatively soluble, and those with multiple charged ions to be less soluble in water. For example, the energy of hydration is greater for Mg^{2+} than for Na^{+}, because Mg^{2+} has a larger charge.

In the series of alkaline earth hydroxides, $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, the lattice energy decreases as the radius of the alkaline earth ion increases (from Mg^{2+} to Ba^{2+}). On the other hand, the energy of hydration is greater for small ions such as Mg^{2+} and is least for large ions such as Ba^{2+}. But the explanation for the observed solubility trend in the alkaline earth hydroxides is that the lattice energy decreases more rapidly in the series $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$ and $Ba(OH)_2$ than does the energy of hydration in the series of Mg^{2+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}. For this reason, the lattice-energy factor dominates this solubility trend, and $Mg(OH)_2$ is less soluble than $Ba(OH)_2$.

Consider the alkaline earth sulphates. Here the lattice energy depends on the sum of the radius of the cations and the radius of the sulphate ion. Because the sulphate ion, SO_4^{2-}, is much larger than the hydroxide ion, OH^{-}, the change in lattice energy in going through the series of sulphates from $MgSO_4$ to $BaSO_4$ is smaller than those of the hydroxides. The lattice energy changes less, and the energy of hydration of the cation decreases by a greater amount. Now the energy of hydration dominates the solubility trend, and the solubility decreases from $MgSO_4$ to $BaSO_4$. Magnesium sulphate is soluble in water, and barium sulphate is very slightly soluble.

Assessment

Assess each student's work throughout section 1.2 (The solution process). To do this, you need to record the performance of each student in your permanent performance list. Your record may be based on how the student:

- participates in discussing Activities 1.2 – 1.5
- participates in presenting opinion after discussion
- performs experiment 1.3
- presents findings of the experiment to the class.
- performs activities given to them between gapped lectures

Give them Exercise 1.3 as class work or homework. Check their work and record their achievements. Also ask the students how intermolecular interaction affects the solubility of liquids in liquids. Ask them what is to be considered in the case of the solubility of solids in liquids. Let the students explain heat of solution and how exothermic and endothermic solutions are formed.

From your record and observations of how the students respond to your questions, check that most students have achieved the competencies suggested for section 1.3. Praise students working above the minimum requirement level and give them extra activities. For these working below the minimum requirement level arrange extra lesson time or give them exercise from the additional questions given in this guide. In addition you may ask the more able students the following questions:

The intermolecular forces in hexane, $C_6 H_{14}$, and heptane, C_7H_{16}, are very similar.

a. Explain why $\Delta H_{solution}$ for the mixing of hexane and heptane is nearly zero.

b. Given that $\Delta H_{solution} = 0$, explain why hexane and heptane mix spontaneously.

Answers:

a. Because both solute and solvent have very similar *London dispersion forces*, then the energy required to separate them individually and the energy released when they are mixed to form a solution are approximately equal.

b. Because there are no strong intermolecular forces that prevent them from mixing.

Additional Questions

* 1. Can a solute dissolve in a given solvent when the strength of the forces between
   a. Solute-solvent particles are comparable to that of solute-solute and solvent-solvent particles,
   b. Solute-solvent particles are weaker than solute-solute and solvent-solvent particles,
   c. Solute-solvent particles are stronger than solute-solute or solvent-solvent particles?

2. Hexane (C_6H_{14}) and methanol (CH_3OH) are miscible as gases but only slightly soluble in each other as liquids. Explain why.

3. What is the relationship between solvation and hydration?

*4. What type of solute-solvent attraction is responsible for the solubility of
   a. Ammonia in water
   b. Sugar in water
   c. Bromine in CCl_4
   d. C_7H_{16} in C_6H_6
   e. $AgNO_3$ in water

* 5. Explain the difference between enthalpy of solution and enthalpy of salvation.

6. Which ion in each pair has the larger heat of hydration?

   a. Na^+ or Cs^+ b. O^{2-} or F^- c. Na^+ or Cl^-

7. The rate of dissolution largely depends on interparticle forces. What other factors influence the rate of dissolution?

Answers to Additional Questions

1. a. Yes   b. No   c. Yes
2. Because in the gaseous state there are enough spaces between molecules. So molecules of C_6H_{14} and CH_3OH are able to mix readily. In the liquid state, the attraction of CH_3OH molecules to each other is stronger than that of CH_3OH and C_6H_{14}. So they are only slightly miscible. In other words, CH_3OH is polar and C_6H_{14} is nonpolar.
3. Solvation can be the same as hydration if the solvent used for the solution process is water. If the solvent is not water, we use the term solvation to describe the process in which solvent molecules surround solute particles.

   a. Hydrogen bonding   b. Hydrogen bonding
   c. London forces   d. London forces.
   e. ion-dipole forces.
4. Enthalpy of solution is the energy change (absorbed or released) when one mole of solute enters solution. Enthalpy of solvation is the energy absorbed or released when solute particles are surrounded by solvent molecules.
5. a. Na^+ b. O^{2-} c. Na^+
6. Temperature, pressure and surface area.

Exercise 1.3

1. a. Rate of dissolution is the speed with which the solute enters into the solution.

   b. Heat of solution is the amount of heat energy absorbed or released when one mole of a substance enters solution.

   c. Solvation is the process by which solute particles are surrounded by solvent molecules.

   d. Hydration energy is the amount of energy released when water molecules surround solute particles.

   e. Lattice energy is the energy required to convert one mole of an ionic solid into isolate gaseous ions.

   f. An ideal solution is a solution for which the heat of solution is zero.
2. a. Ion-dipole   b. Ion-dipole   c. London forces
3. Because the lattice energy of calcium carbonate $CaCO_3$ is higher than that of Na_2CO_3. Thus, $CaCO_3$ is less soluble than Na_2CO_3.

1.4 SOLUBILITY AS AN EQUILIBRIUM PROCESS

Period Allotted: 4 periods

Competencies

At the end of this section, students will be able to:

- *describe the distinctions among unsaturated, saturated and supersaturated solutions;*
- *prepare a supersaturated solution of sodium phosphate;*
- *explain the equilibrium nature of a saturated solution;*
- *define solubility and describe the factors that affect the solubility of substances;*
- *investigate the effect of temperature on the solubility of sodium sulphate;*
- *conduct an experiment to determine the solubility of table salt and sugar;*
- *state Henry's law; and*
- *use Henry's law to calculate the concentration of gaseous solute in a solution.*

Forward Planning

Thoroughly read the contents on solubility as an equilibrium process (section 1.4) from the student's text and other reference materials. Prepare a plan of your own that shows the contents, Activities and experiment you are going to deal with during each period. The plan must ensure that the entire content of the section will be covered within four periods. In your plan, indicate the duration of time allotted to every activity you perform in each period, such as group discussion, presentation, gapped lecture etc. Read the teacher's guide to get information about the methodologies you implement for the section and to gain more ideas about Activities 1.6 –1.8.

Prepare chemicals, apparatus and materials required to perform experiment 1.4. Carry out the experiment yourself before you allow students to do it.

Teaching Aids

Refer to the students' text for materials required to perform Experiment 1.4.

Subject Matter Presentation

Use group discussion, experiment, question-and-answer methodologies to teach the contents in this part.

Be sure that the students understand that solutions are mixtures but are not pure compounds. They are not governed by the law of constant composition. Hence, one can make solutions of different concentrations from the same solute and solvent. So, solutions can be unsaturated, saturated or supersaturated.

Start the lesson with Activity 1.6. The activity is designed to help students realize the difference between saturated and unsaturated solutions. So, let the students do Activity 1.6. To do this, group them and encourage them to discuss the questions. In your harmonizing lecture, let your students understand the following facts.

A water-soluble substance dissolves in water until no more can dissolve at a given temperature. At the start, the solution is unsaturated. It can dissolve more solute. As time passes, more solute continues to dissolve into the solution, which means that the concentration increases. Eventually, the state of dynamic equilibrium will be reached, in which the excess undissolved solute is in equilibrium with the dissolved solute. This solution is called saturated solution. It contains the maximum amount of dissolved solute at a given temperature in the presence of undissolved solute.

Let them perform Experiment 1.4 in groups and discover the amount of NaCl that dissolves per 100 grams of water at a given temperature.

An **unsaturated solution** is a solution that is not in equilibrium in it with the dissolved substance. Such a solution has room for more solute to dissolve. In some cases, we can prepare a solution that contains more than the equilibrium amount of dissolved solute. Such a solution is called **supersaturated solution**. Activity 1.7 shows this phenomenon.

From the activity, make sure that students have realized what a supersaturated solution is. Tell them that when a supersaturated solution is cooled and stirred, the excess solute dissolved will settle down. When a "seed" crystal of the solute is added, it increases in size due to crystallization of the excess salt on the surface of the seed.

Effect of Temperature on Solubility

Use group discussion and gapped lecture as methodologies for this topic.

As you proceed with the lesson **Effect of Temperature on Solubility**, start with Activity 1.8. The activity enables students to discover whether or not solubility of substances depends on temperature.

Let your students perform activity 1.8 for a few minutes. In your explanation that follows the students' observation and reflection, let the students notice that the solubility of a substance at a given temperature is the maximum quantity of substance, in moles or grams, which dissolves in a solvent to form one liter of saturated solution at that temperature. The solubility of substances changes with temperature.

Help the students to remember the factors that affect the rates of chemical reactions and draw parallel reasoning with the factors that affect the rates of dissolution. Among the many factors that determine the solubility, temperature should be given special consideration.

Let the students notice that the conventional reference for solubility is the number of grams of solute that can dissolve in 100 g of solvent at a particular temperature. Let them refer to the solubility curve given in their textbook. The curve shows that the solubility of most salts increases with increasing temperature. Help them to read the solubility of each salt given in the solubility curve at different temperature. For example, a saturated KCl solution at 10°C will have 31 g of KCl dissolved in 100 g of water. Raising the temperature of the mixture to 30°C will increase the amount of dissolved KCl to 37 g.

The students should appreciate that the solubilities of almost all solid solutes increases with temperature, whereas those of gases invariably decrease. Let the students form up into groups and then discuss in for a few minutes why the solubility of gases decreases with an increase in temperature. Next have some groups present their opinions.

After the presentations, tell them that the dissolution of gases is an exothermic process. So for most exothermic solution processes, solubility decreases with an increase in temperature.

Effect of Pressure on Solubility

Gapped lecture is suggested as a methodology to teach this topic. As you proceed to discuss *The Effect of Pressure on Solubility*, let your students be aware of the fact that the solubility of gaseous solutes depends, in part, on pressure. This dependence is expressed with **Henry's law,** which relates the concentration of a dissolved gas, C_g, to its partial pressure, P_g, over the solution.

$C_g = k_g P_g$, where k_g is Henry's law constant.

Assessment

To assess each student’s work throughout section 1.4, watch carefully how every student is doing and record the performance in the permanent list. You can do this by making a record of each student’s participation:

- in discussing Activities 1.6 – 1.8
- in presenting the findings of the group after discussion
- in performing experiment 1.4 and presenting the results of the experiment to the rest of the class
- in answering questions raised during gapped lectures and activities given to them between the lectures

Give them exercise 1.4 as homework. Check their work and record their achievements.

Ask the students to define terms like solubility, unsaturated solution, saturated solution and supersaturated solution. Ask them to explain the effect of temperature on the solubility of solid solutes and that of pressure on the solubility of gases.

Based on the records you have and their attempts in answering the oral questions, check whether or not the competencies suggested for section 1.4 are achieved by most of the students. Praise those working above the minimum requirement level and give them extra activities. With regard to students working below the minimum requirement level, arrange extra lesson time or give them additional exercise.

Additional Questions

* 1. You are given three bottles containing aqueous solutions of X one saturated, one unsaturated and one supersaturated. How can you identify which solution is which?

* 2. The solubilities of $AgNO_3$ at 10°C and 20°C are 170g and 222g per 100g H_2O. What is the sign of ΔH_{soln} for $AgNO_3$?

3. The heat of solution of NaI is −7.5 kJ/mol. Is NaI more soluble at 20°C or 60°C?
4. What is the effect of pressure on the solubility of solid and liquid solutes in water?

* 5. When a water sample in a beaker is heated, bubbles escape out of the water. What does this indicate?

* 6. When we open Ambo mineral water or soft drinks, we hear a hissing sound and also observe bubbles escaping out of the water or drink. What does the escape of the gas bubbles indicate?

Answers to Additional Questions

1. By adding a crystal of X in each solution. The size of the crystal decreases, remains the same and increases when added into unsaturated, saturated and super saturated solutions, respectively.
2. ΔH_{soln} is positive.
3. More soluble at 20°C.
4. It has almost no effect.
5. Solubility of gases decreases with an increase in temperature. The bubbles that escape out of the water are bubbles of oxygen gas.
6. Solubility of gases in a solvent decreases with a decrease in the partial pressure of the gas.

Answers to Exercise 1.4

1. 5×10^{-4} M
2. 9.3×10^{-6} M

1.5 WAYS OF EXPRESSING CONCENTRATIONS OF SOLUTION

Periods Allotted: 7 periods

Competencies

At the end of this section, students will be able to:

- *define terms like solution, mass percentages, ppm and ppb of a solute in a solution;*
- *calculate the mass percentage, ppm and ppb of a solute in a solution from given information;*
- *define mole fraction, molarity, equivalent mass, number of equivalents, normality and molality;*
- *calculate mole fraction, molarity, normality and molality of a solution from given information;*
- *interconvert various concentration expressions.*

Forward Planning

You should familiarize yourself with the concentration units mentioned in the students' textbook (mass percentage, ppm and ppb, mole fraction, molarity, normality and molality). Prepare a plan of your own that shows the topics you are going to deal with during each period in such a way to cover the whole content of the section within seven periods. Also indicate the duration of time you allot for gapped lectures, class works, stabilization and other activities you perform during each period in your plan.

Subject Matter Presentation

You may use gapped lectures combined with students' participation by doing class work and home work using the exercises suggested. You may prepare short activities that will help to discuss some of the subtopics.

You may start your presentation of this section by mentioning that the quantity of solute in a solution is a very important property of a solution. Then define concentration as the quantity of solute per unit quantity of solution.

Let the students know that there are many ways in which the concentration of a solution can be described and defined in concentration units.

Mass percentage ppm, and ppb of solute

Using examples explain mass percentage of solute in a solution and then introduce ppm and ppb as units used to express the concentration of solutions containing very small amounts of solutes. Let the students practice calculating mass percentage ppm and ppb. You can give Exercise 1.5 as class work or homework.

Mole Fraction (X)

Explain that the mole fraction refers to the ratio of the number of moles of one substance to the total number of moles of all substances in the solution.

In a solution containing n_A moles of solute and n_B moles of solvent, the mole fraction of solute, x_A, and the mole fraction of the solvent, x_B, can be expressed as follows:

$$x_A = \frac{n_A}{n_A + n_B} \qquad x_B = \frac{n_B}{n_A + n_B}$$

Let the students know that the sum of the mole fractions of all the components in a solution must be equal to **one**. Give them Exercise 1.6 as class work or homework so that they can practice calculating mole fractions.

Molarity (M)

Illustrative examples on how to calculate the molarities of solutions are given in the students' textbook. Students should know how the problems are solved and understand each step in the conversion process. You may add some more examples and encourage the students to do the given exercises in the textbook.

Normality (N)

The students should be familiar with normality (N), which is defined as the number of equivalents of solute per liter of solution. Note that determining the number of equivalents is not as simple as calculating the number of moles. To find the number of equivalents, one should know the equivalent mass.

Note for the Teacher

The equivalent mass of an acid is that fraction of the molar mass which contains, or can supply for reaction, one mole of H^+. In other words, the equivalent mass is the molar mass, divided by the number of H^+ furnished per molecule. For example, the equivalent mass of HCl is equal to its molar mass because it contains only one acidic hydrogen per molecule. The equivalent mass of H_2SO_4 is half of its molar mass because it contains two acidic hydrogens per molecule. In a similar way, the equivalent mass of a base is its molar mass, divided by the number of OH^- furnished per molecule. The equivalent mass of NaOH is equal to its molar mass, while the equivalent mass of $Al(OH)_3$ is its molar mass divided by three, because it furnishes three OH^- ions.

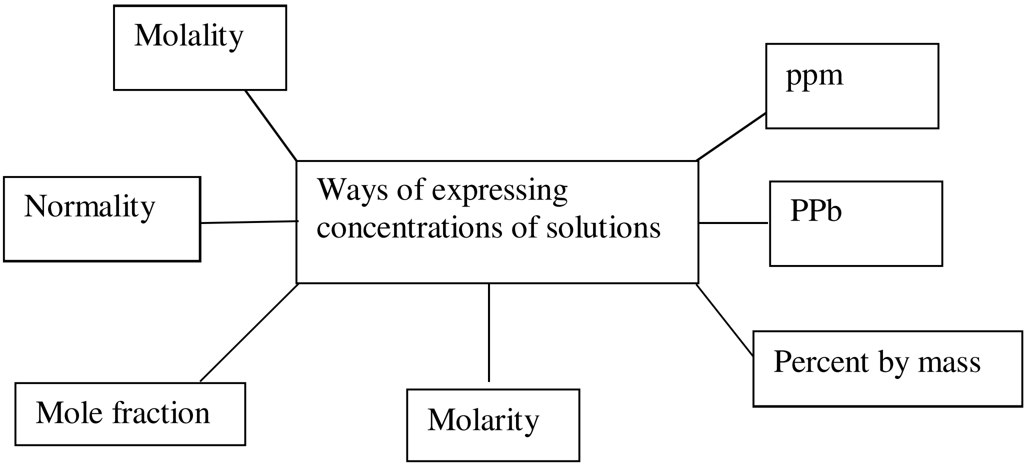
The equivalent mass of an oxidizing agent or reducing agent for a particular reaction is equal to its molar mass, divided by the total number of moles of electrons gained or lost when the reaction of one mole occurs. You may have a look at the illustrative example given in the student's textbook concerning this topic. Let the students practice calculations involving normality.

Molality (m)

Molality is defined as the number of moles of solute per kilogram of solvent. A solution concentration that compares moles of solute and kilograms of solvent (molality) is most useful in those calculations dealing with colligative properties of solutions, which will be studied in section 1.9 of this unit.

Students should also learn how to convert among the different concentration units.

As a summary of ways of expressing concentration of solution let students draw a spider diagram. See whether or not their diagrams look like the following



Assessment

To assess each student’s work in section 1.5 (ways of expressing concentrations of solutions), record each student’s performance with exercises. To do this, give them exercises 1.5 – 1.10, correct their work and record their achievements.

Give as many exercises as possible in order to have the students practice. The exercises you give them should consider the slow learners and gifted students. For example, for slow learners, you may give direct questions such as giving the students the mass of a solute, and the volume of a solution and then asking the molarity. For gifted students, you may give the density of the solution instead of giving the mass of the solute.

From your records, make sure that students working at the minimum requirement level will be able to: calculate the mass percentage, ppm and ppb of a solute in a solution, mole fraction of a solute and a solvent in a solution, molarity of a solution, normality of a solution, molality of a solution from given information and also be able to interconvert concentration units.

Additional Questions

* 1. Wastewater from a cement factory contains 0.22 g Ca^{+2} ion and 0.066 g Mg^{2+} ion per 100L solution. The density of the solution is 1.001 g/mL. Calculate the Ca^{2+} and Mg^{2+} concentrations in ppm (by mass).

2. Gold occurs in a sample of sea water at a concentration of 1.1×10^{-2} ppb. How many litres of sea water must be processed to recover 31.1 g of gold? (Density of sea water = 1.025g/mL.)

3. An 8.0% by-mass aqueous solution of ammonia has a density of 0.9651g/mL. Calculate the molality, molarity and mole fraction of ammonia.

*4. What is the percent by mass of the solute in each of the following?
   a. 5.50 g of NaBr is 78.2 g of solution
   b. 4.5 g of tolnene in 29 g of benzene

5. 42.6 g of $K_2Cr_2O_7$is dissolved in enough dilute HCl to make up exactly one litre of solution. Calculate the molarity and normality of the solution when $Cr_2O_7^{2-}$ reacts according to the following reaction.

$$Cr_2O_7^{2-} + Cl^- \rightarrow Cr^{+3} + Cl_2 + H_2O$$

Answers to Additional Questions

1. 2.2 ppm Ca^{2+}; 0.66 ppm Mg^{2+}
2. 4.4×10^9 L.
3. 5.13m; 4.54M; $X_{NH_3} = 0.0845$
4. a. 7.03% b. 13.0%
5. 0.87 N.

Answers to Exercises

1.5 1. 1.61%

2. a 800 ppm b. 800,000 ppb

1.6 1. a. X H_2O = 0.75% b. X alcohol = 0.25%

2. a. $X_{(glycerine)} = 0.20$ b. $X_{(water)} = 0.80$

3. $X_{Na^+} = X_{Cl^-} = 0.0098$

1.7 1. a. 0.21 M b 2.2%

2. By dissolving 3.77 g KNO_3 in enough water until the volume of the solution is 250 mL.

1.8 1. a. 2 equivalents of solute b. 0.5 equivalents of solute

c. 0.1 equivalents of solute

2. 6.84 g
3. 0.203 M, 1.22 N

1.9 a. $X_{HCl} = 0.22$ b. 15 m

c. We need to know the volume of solution. If we were given the density of the solution, we could convert the mass of the solution to the volume of the solution and then calculate the molarity.

1.10 1. a. % HCl = 36.2

b. molality of HCl = 15.5 m

c. $X_{HCl} = 0.219$

2. a. 23.8m

1.6 PREPARATION OF SOLUTIONS

Period Allotted: 2 periods

Competencies

At the end of this section, students will be able to:

- *prepare molar solutions, normal solutions and molal solutions of different substances;*
- *explain the dilution process and calculate the volume or concentration changes during dilution of a solution;*
- *prepare a dilute solution from a concentrated solution.*

Forward Planning

Make a special preparation to start this topic, since there is an experiment to show how solution is formed. Arrange the necessary materials needed for Experiments 1.5 and 1.6. Since there is no danger to doing this experiment, let the students do the experiment by themselves.

Teaching Aids

Refer to the student's text for the materials required to perform Experiments 1.5 and 1.6.

Subject Matter Presentation

Use experiment (as the main methodology) and a mini-lecture for this part.

A common task in school, medical, industrial and other chemical laboratories is the preparation of solutions of known concentration. For aqueous solutions, distilled, demineralized or deionized water is used. Solutions are usually prepared from liquids or solids, and occasionally gases, as solutes.

Let the student perform **Experiment 1.5** to prepare a solution of known concentration. Help them to do this experiment correctly.

After they complete Experiment 1.5, make sure that their results in the observation and analysis part is as follows.

For procedure 1

a. 0.2 mole NaCl b. 1.0M and 1.0 N

For procedure 2

a. 25.48% sugar b. $x_{sugar} = 0.018, x_{H_2O} = 0.982$ c. 1.0 m

In addition to the experiment given, the students may prepare different solutions of known concentration using chemicals that are available in the school laboratory. To do this, they need to know the mass of the solute and the volume of the solution that they have prepared (in case of a molal solution, the weight of the solvent). Encourage the students to prepare solutions of different concentrations.

Diluting Solutions

Use group discussion, gapped lectures and experiment as methodologies for this topic. Before you continue with dilution, let students do activity 1.9 in groups for a few minutes and discover the effect of dilution on concentration. Then have some groups present their conclusions. To harmonize concepts, tell them that the colour intensity of solutions of $CuSO_4$ and $KMnO_4$ decreases. The decrease in colour intensity of the solutions is due to a decrease in concentrations of $CuSO_4$ and $KMnO_4$ as we dilute them to 50 mL. The same thing is true for colourless solutions.

Making dilutions is another important application of concentration units. **Experiment 1.6** helps the students to understand how to prepare a solution of lower concentration from a stock solution. A solution can be made less concentrated by dilution with solvent. The number of moles of solute does not change when more solvent is added to the solution. If a solution is diluted from V_i to V_f, the concentration of that solution changes according to the equation:

$$C_i V_i = C_f V_f$$

Let the students know that stock solutions are very concentrated. To use them, they usually need to be diluted, using the dilution equation given above.

Use the following note for the Observations and Analysis part of Experiment 1.6.

a. The addition of water to concentrated acid produces a great deal of heat, which causes spattering of the acid. Thus concentrated acid is added to water, since the water absorbs the heat produced so that the acid doesn't spatter.

b. In the experiment, the mass and number of moles of the acid remains the same. But, the volume of the solution increased.

c. 3.0 M d. 0.18 M

Assessment

To assess each student's work in relation to the preparation of solutions (section 1.6), record how each student participates in

- discussing Activity 1.9
- presenting the opinion of the group to the rest of the class
- performing experiments 1.5 and 1.6
- presenting the observation of the group resending the experiments to the rest of the class

Give them exercise 1.11 as homework. Check their work and record their achievements.

In addition to the exercises given in the students' text book in this section, give as many exercises as possible to the students and correct their work. From your record, make sure that students working at the minimum requirement level are able to: prepare molar solutions, normal solutions and molal solutions of different substances; explain dilution process and calculate the volume or concentration changes during dilution of solution; and prepare a dilute solution from concentrated solution. Appreciate students working above the minimum requirement level. Assist those working below the minimum requirement level by arranging extra lesson time.

Additional Questions

1. How do you prepare
   a. 1.0 L of a 3.5 M solution of H_2SO_4
   b. 500 mL of a 0.25 M solution of KOH
2. How can you prepare a 2.4% by-mass KOH solution in 8.5 g of solvent?

*3. 200 mL of 1 2.0 M HNO_3 is prepared from a stock solution of 16.0 M HNO_3.
   a. What volume of stock solution of 16.0 M HNO_3 is used?
   b. What volume of water should be added to the stock solution?

Answers to Additional Questions

1. a. By dissolving 343 g of H_2SO_4 b. By dissolving 7.0 g KOH
2. By dissolving 0.209 g KOH
3. a. 25 ml of 16.0M HNO_3 b. 175 ml of water

Answer to Exercise 1.11

a. We must dilute 3 mL of the 4 M HNO_3 with sufficient water to give the final volume of 60 mL.

b. 400 mL of water must be added.

1.7 SOLUTION STOICHIOMETRY

Periods Allotted: 3 periods

Competencies

At the end of this section, students will be able to:

- *use stoichiometerically equivalent molar ratios to calculate the amounts of reactants and products in a reaction of pure and dissolved substances.*

Forward Planning

Read about solution stoichiomentry (section 1.7) to review the stoichiometry of reactions of pure substances and apply the rules to the substances in solution. Plan which content and activity you will treat in each period so that the entire content can be covered with in three periods

Subject Matter Presentation

Use group discussion and gapped lecture as methodologies for this section.
Start teaching this section by using activity 1.10. The activity enables them to recall what stoichiometry is and also understand its application to solutions. So, let the students discuss Activity 1.10 in groups for a few minutes and have some groups present their opinion to the class. After the presentations, harmonize their discussion by adding more points. Help your students to know the principles of stiochiometry that they have studied in Grade 9 and apply them to reactions in solution as well.

Let your students notice that the mole concept allows us to use the quantitative information available in a balanced chemical equation. So, as in any other reaction, students should write balanced chemical equations to describe the chemical reaction.

Let them recall that the coefficients in a balanced equation give the relative number of moles of reactants and products. To use this information, we have to convert the amount of substances involved in a reaction into moles. When we are dealing with masses of a substance, we use molar mass to achieve this conversion. When we are working with solutions of known concentration, however, we use the concentration and volume of the solution to determine the number of moles.

Number of moles of solute = Concentration of solution x Volume of the solution

For example, what is the number of moles of H_2O that can be formed when 25.0 mL of 0.100 M HNO_3 solution is completely neutralized by NaOH?

The product of the molar concentration of the solution and its volume in liters give the number of moles of solute (HNO_3).

Number of moles of HNO_3 = M × L

$$= 0.100 \frac{\text{mol}}{\text{L}} \times 0.025\ \text{L} = 2.5 \times 10^{-3} \text{ moles of } HNO_3$$

Then we write the neutalization reaction and give the balanced chemical equation for the reaction.

$$HNO_3(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaNO_3(aq)$$

Thus, 1 mole of HNO_3 produces 1 mole H_2O.

Therefore,

Number of moles of H_2O produced from 2.5×10^{-3} moles of HNO_3 =

$$\frac{2.5 \times 10^{-3} \text{ mole of } HNO_3 \times 1 \text{ mole of } H_2O}{1 \text{ mole of } HNO_3} = 2.5 \times 10^{-3} \text{ moles of } H_2O$$

Let the students calculate amounts of reactants and products in a reaction of pure substances and in a reaction of dissolved substances (solution), using the relations between mole – mass, mole – volume and mole – number of particles.

Assessment

To assess each student's work throughout the teaching learning process on solution stoichiometry, observe and record how student participates in discussing Activity 1.10, in presenting opinions of the group, and in doing activities given between gapped lectures.

Ask the students how to get the number of mole in stoichiometric calculations of pure substances and substances in solutions. Encourage both the slow learners and gifted students to do Exercise 1.12. Correct their work and record their achievements. From the cumulative record you have, make sure that students working at the minimum requirement level are able to use stoichiometrically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substances. Encourage students working above the minimum requirement level to continue woring hard. For those working below the minimum requirement level, arrange extra lesson time or give them exercises.

Additional Questions

*1. What mass of each product results if 750 mL of 6.0M H_3PO_4 reacts according to the following equation?

$$H_3PO_4 + Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$$

2. How many milliliters of 18.0M H_2SO_4 would be required to react with 250 mL of 2.5 M Al $(OH)_3$?

   $H_2SO_4 + Al(OH)_3 \rightarrow Al_2(SO_4)_3 + H_2O$

3. If 75 mL of a $AgNO_3$ solution reacts with enough Cu to produce 0.25 g of Ag by single displacement, what is the molarity of the initial $AgNO_3$ solution?

   $AgNO_3 + Cu \rightarrow Cu(NO_3)_2 + Ag$

*4. What mass of NaOH should be contained in 250 mL solution to neutralize completely 500 mL of a 0.2 M H_2SO_4, according to the following equation?

   $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

Answers to Additional Questions

1. 698 g $Ca_3(PO_4)_2$ and 243 g H_2O.
2. 52.1 ml of 18.0 M H_2SO_4.
3. 0.0309 M
4. 8.0 g NaOH

Answer to Exercise 1.12

1. 0.5L
2. 50mL
3. 75mL

1.8 DESCRIBING REACTIONS IN SOLUTION

Periods Allotted: 1 period

Competencies

At the end of this section, students will be able to:

- *explain the relationship between reacting ions, spectator ions, precipitation and solubility;*
- *write net ionic equations.*

Forward planning

Prepare yourself to present this section in such a way that reaction equations in solutions may be molecular or ionic. To write ionic equations, it is important to know which substances can be dissolved in water to form aqueous solutions and which cannot dissolve in water. So, review this part before you present this section.

Subject Matter Presentation

Use group discussion and gapped lecture to teach this section. Let the students define molecular equation from what they have learned in Grade 8. Then let them differentiate between molecular equation and ionic equation.

In writing chemical equations for reactions in solution, it is often useful to indicate explicitly whether the dissolved substances are present predominantly as ions or molecules. Activity 1.11 is suggested to help students identify the form in which a dissolved substance exists in aqueous solution. So, let the students discuss Activity 1.11.

Make sure that they have included the following points in their discussion. Calcium hydroxide, $Ca(OH)_2$, and sodium carbonate, Na_2CO_3, are both soluble ionic substances and are strong electrolytes. When they dissolve in water, they go into solution as ions. Each formula unit of $Ca(OH)_2$ forms one Ca^{2+} ion and two OH^- ions in solution. So, instead of $Ca(OH)_2(aq)$, it would be better to write $Ca^{2+}(aq) + 2OH^-(aq)$. Similarly, each formula unit of Na_2CO_3 forms two Na^+ ions and one CO_3^{2-} ion in solution, and this can be written as $2Na^+(aq) + CO_3^{2-}(aq)$ instead of as $Na_2CO_3(aq)$. The reactant side of the reaction becomes

$$Ca^{2+}(aq) + 2OH^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \longrightarrow$$

On the product side of the equation, $CaCO_3(s)$ is an ionic compound, but the ions are fixed in particular sites in the crystalline solid. We leave the formula as $CaCO_3(s)$ to convey this information in the equation. On the other hand, NaOH is a soluble ionic compound and is a strong electrolyte. Also, it dissolves in aqueous solution to give freely moving ions. Therefore, we can write it as $Na^+(aq) + OH^-(aq)$.

The complete equation is

$$Ca^{2+}(aq) + 2OH^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s) + 2Na+(aq) + 2OH^-(aq)$$

The ions appearing on both sides of the equation (Na^+ and OH^-) are called **spectator ions,** as they do not take part in the reaction. They can be cancelled on both sides to express the essential reaction that occurs.

$$Ca^{2+}(aq) + \cancel{2OH^-}(aq) + \cancel{2Na^+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s) + \cancel{2Na^+}(aq) + \cancel{2OH^-}(aq)$$

The resulting equation is:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

This is the **net ionic equation**, which is an ionic equation without the spectator ions. The reaction that actually occurs at an ionic level is that between calcium ions and carbonate ions, which forms solid calcium carbonate.

After you introduce molecular, ionic and net ionic equations, write the following equation on the blackboard.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$$

Let the students discuss, in groups, for a few minutes how to:

a. write the ionic equation

b. write the net ionic equation

c. identify the spectator ion(s)

d. identify the nature of the substances whose formulas appear in molecular form in both the ionic and net ionic equation

When the discussion is over, let students from two different groups suggest their opinions to the class. Following the presentations, tell them the following

a. The ionic equation is

$$CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + H_2O(\ell) + CO_2(g)$$

b. The net ionic equation is

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$$

c. The spectator ion is Cl^-

d. The substances whose formulas appear in molecular form in ionic and net ionic equations are solids, pure liquids, gases and weak electrolytes.

Assessment

To assess each student's work related to section 1.8, use exercises 1.12. Molecular equations are given and the students are asked to write the net ionic equation and to identify the spectator ions. Encourage the students to do these exercises and give them additional work. Correct their work. Based on their achievements, be sure that students working at the minimum requirement level are able to explain the relationships between reacting ions, spectator ions, precipitation and solubility and to write net ionic equations.

Additional Questions

1. Write the net ionic equation for each of the following reactions.

a. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,NaCl(aq)$

b. $(NH_4)_2CO_3(aq) + CaCl_2(aq) \rightarrow 2NH_4Cl(aq) + CaCO_3(s)$

c. $Pb(NO_3)_2(aq) + CuSO_4(aq) \rightarrow PbSO_4(s) + Cu(NO_3)_2(aq)$

d. $Mg(HCO_3)_2(aq) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2CO_2(g) + 2H_2O(\ell)$

Answers to Additional Questions

1. a. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

b. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

c. $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

d. $2HCO_3^-(ag) + 2H^+(ag) \rightarrow 2CO_2(s) + 2H_2O(\ell)$

Answers to Exercise 1.13

a. $2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)$

Spectator ions – Na^+ and NO_3^-

b. $2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$

Spectator ions – Ca^{2+} and ClO_4^-

c. $2H^+(aq) + Mg(OH)_2(s) \rightarrow 2H_2O(l) + Mg^{2+}(aq)$

Spectator ion – NO_3^-

d. $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

Spectator ions – Na^+ and NO_3^-

1.9 COLLIGATIVE PROPERTIES OF SOLUTIONS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *define colligative properties;*
- *list the important properties of solvents that are affected by the formation of a solution;*
- *state Raoult's law;*
- *explain the change in the colligative properties of a solvent when a certain solute is added;*
- *describe van't Hoff's factor;*
- *calculate the vapor pressure, boiling point, and freezing point of a solvent after a certain amount of solute is dissolved in it;*
- *define osmosis and osmotic pressure of a solution;*
- *calculate the osmotic pressure of a solution;*
- *compare and contrast changes in the colligative properties of electrolytic and nonelectrolytic solutions.*

Forward planning

Get ready to present this section by studying the four physical properties of solutions that depend only on the number of solute particles. These properties are known as colligative properties. They are: vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure. Design a plan that shows the topics and activities you are going to treat during each period in such a manner that the entire contents will be covered within five periods.

Subject matter presentation

Encourage the students to answer the in-text question as you start. In your mini-lecture that follows, harmonize their responses by giving them the exact answers to the question.

Students should appreciate that solutions have different physical properties from those of the pure solvents. These differences in properties depend not on the nature of the solute dissolved in the solvent, but on the number of particles per unit volume of the solution. *Hence, they are called colligative, because the word colligative means collective.*

Vapour pressure lowering

Use the question-and-answer methodology to teach this section.

Solutes can affect the properties of a solvent. For example, the water in syrup evaporates much less quickly than pure water, because the vapor pressure of water in the syrup is lower than the vapor pressure of pure water. The presence of the solute sugar in the syrup solution is responsible for the difference.

Ask them to explain how dissolved solute can lower the vapor pressure of solution. After their responses, tell them that solute particles take positions originally occupied by solvent particles. This lowers the number of solvent particles that are closer to the surface and entering the gaseous state. This situation in turn reduces the vapor pressure of the solution. Then continue with **Raoult's law.** Raoult's law states that the vapor pressure of any component of an ideal solution is equal to the vapor pressure of the pure component, multiplied by its mole fraction in the solution. For the vapor pressure of the solvent in a solution of a nonvolatile solute, the mathematical form of this law is

$$P_{solvent} = X_{solvent} \; P^{o}_{solvent}$$

Where $P_{solvent}$ is the vapor pressure of the solvent in the solution, $X_{solvent}$ is the mole fraction of the solvent in the solution, and $P^{o}_{solvent}$ is the vapor pressure of the pure solvent.

In a solution containing only one solute, $X_{solvent} = 1 - X_{solute}$, Raoult's law equation can be written as:

$$P_{solvent} = (1 - X_{solute})\, P^{o}_{solvent}$$

$$P_{solvent} = P^{o}_{solvent} - P^{o}_{solvent}\, X_{solute}$$

$$P^{o}_{Solvent} - P_{solvent} = P^{o}_{solvent}\, X_{solute}$$

$$\Delta P = P^{o}_{solvent}\, X_{Solute}$$

After you derived the mathematical expression for ΔP, ask the students what they understand from the relationship. After their responses, have them know that the decrease in vapor pressure, ΔP, is directly proportional to the solute concentration, measured in mole fraction. Support your explanation using examples and use Exercise 1.14 appropriately. Introduce them to the characteristics of ideal solutions. Also, give them information about the situations in which nonideal solutions can have negative or positive deviations from Raoult's law.

Boiling point Elevation

Use gapped lecture to teach this topic. Begin presenting *boiling point elevation* by using the in-text question. In your explanation, let your students notice that the presence of a nonvolatile solute lowers the vapor pressure of a solvent and, therefore, affects the boiling point of the solution. The boiling point elevation (ΔT_b) is defined as the boiling point of the solution (T_b) minus the boiling point of the pure solvent (T^{o}_{b}):

$$\Delta T_b = T_b - T^{o}_{b}$$

The value of ΔT_b is directly proportional to the concentration, in molality, of the solution. That is

$$\boxed{\Delta T_b = K_b m}$$

where m is the molality of the solution, and K_b is the molal boiling point elevation constant. The unit of K_b is $^{o}C/m$. Support your explanation, using examples. You may give Exercise 1.15 either as homework or class work.

Freezing point depression

Use group discussion and gapped lecture as methodologies to teach this topic.

As you proceed to the lesson on freezing *point depression*, start with Activity 1.12. The activity helps the students discover how the presence of a solute influences the freezing point of a solvent. Start by asking your students to discuss the issue in Activity 1.12. As you facilitate their discussion, record their doubts. Let some of the students change groups and share the opinion of their previous group with others. In your mini-lecture that follows, harmonize their ideas with the exact concept presented in the text.

Emphasize the following points during the teaching-learning process. The addition of a solute affects the vapor pressure of a liquid solvent. Since changes of state depend on vapor pressure, the presence of a solute also affects the freezing point and boiling point of a solvent. Thus, with the addition of a solute to a solvent, the freezing point of a solution is lower than the freezing point of pure solvent.

For example, water freezes at 0^oC at a pressure of 1 atm. At that temperature, the vapor pressures of ice and water are the same. But, when the solute is added, the vapor pressure of the liquid solution is less than that of ice and thus it does not freeze, as the freezing point is depressed.

As well as understanding the actual effect of the added solute, the students should master calculations related to freezing point lowering. This dependence is given quantitatively as

$\Delta T_f = K_f m$ where ΔT_f and m is concentration molaity.

From their discussions, the students may observe that the addition of ethylene glycol to water decreases the freezing point of the solution and increases the boiling point of water (Activity 1.12). Salts like NaCl and $CaCl_2$ are used to de-ice roads because they cause the ice to melt by lowering the freezing point and lowering its vapor pressure. Let students do Exercise 1.16 as class work or homework.

Osmosis and osmotic pressure

Gapped lecture and group discussion are the suggested methodologies to teach this topic.

As you move to the lesson on *Osmosis and osmotic pressure*, you may start with the suggested in-text question. You may also prepare an activity or other questions to be discussed by students. For example, you may ask them to define osmosis and osmotic pressure. After their responses, in your lecture that follows, explain to the students that the osmotic pressure of a solution is different from that of a pure solvent. If we put a solution and a pure solvent next to each other, joined by a semi-permeable membrane that allows a solvent to pass through it but not solute molecules, then solvent molecules will preferentially "pass" from the pure solvent into the solution until equilibrium is attained. This is due to the rising pressure on the solution side as it keeps gaining solvent molecules.

Tell the students the formula for the extra pressure needed on the solution side to stop the flow of solvent, which is called osmotic pressure, π.

$$\pi = MRT$$

where π is osmotic pressure, M is the molarity of the solution, R is gas constant and T is absolute temperature. Support your explanation, using examples given in the text or from your own. You may give Exercise 1.17 as homework.

Before you conclude your presentation of osmotic pressure, let students discuss Activity 1.13. The activity is designed to help students realize how the diffusion of water proceeds from a less concentrated solution to a more concentrated one. So, let the students do activity 1.13 in groups. After they complete it, let some groups present their findings to the class. Next, make sure the students discovered that the water level in the potato kept in 20% sucrose solution decreases to a greater extent compared to the others. The water level kept in the potato in 5% sucrose solution decreases only slightly, compared to that of the others.

Colligative Properties of Electrolytic Solutions

You are advised to implement group discussion and gapped lecture as methodologies for this topic.

The final topic of this subsection of this topic is *Colligative Properties of Electrolytic Solutions.* Before you deal with colligative properties of electrolyte solutions, use Activity 1.14 to start the lesson. The activity is designed to enable students to differentiate the effects of electrolytes and nonelectrolytes on the properties of solutions. Therefore, let students discuss activity 1.14 in groups for a few minutes and have some groups present their opinions to the class.

After the presentation, tell them the difference between electrolytes and non-electrolytes. In addition to this, let them know that electrolytes produce greater colligative effects than do nonelectrolytes. This is because electrolytes produce more particles per mole than do nonelectrolytes. For example, one mole of glucose can produce one mole of glucose molecules on dissolution. However, one mole of $CaCl_2$ produces approximately one mole of Ca^{2+} and two moles of Cl^- ions and a total of three moles of solute particles. Thus, one mole of $CaCl_2$ produces more colligative effects than does one mole of glucose. Based on this example, ask the students to explain the difference between the effect of nonelectrolyte solutes and electrolyte solutes on colligative properties. Then inform to them that the electrolytes dissociate into ions in electrolyse solution, and that one unit of an electrolyte compound separates into two or more particles when it dissolves. Here it is important to remember that it is the number of solute particles that determines the colligative properties of a solution. For example, each unit of NaCl dissociates into ions, Na^+ and Cl^-. Thus, the colligative properties of a 0.1 m NaCl solution should be twice as great as that of 0.1 m solution containing a nonelectrolyte. To account for this effect, we define a quantity known as the Van't Hoff factor (i).

$$i = \frac{\text{measured value for electrolyte solute}}{\text{expected value for nonelectrolyte solution}}$$

To calculate the colligative properties of electrolyte solutions, we incorporate the Van't Hoff factor into the above equations. Thus

For vapor pressure lowering: $\Delta P = i\left(X_{\text{solute}}\ P^{o}_{\text{solvent}}\right)$

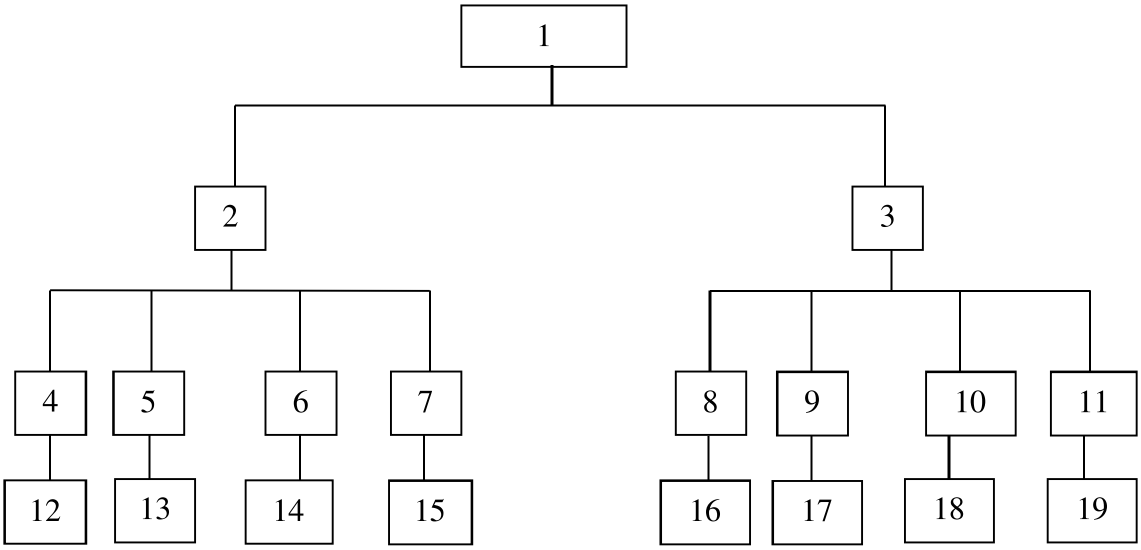
For boiling point elevation: $\Delta T_b = i\left(k_b m\right)$

For freezing point depression: $\Delta T_f = i\left(k_f m\right)$

For osmotic pressure: $\pi = i\,(MRT)$

Finally, summarize the chapter and encourage students to do the review exercises. As a summary of colligative properties of solutions, let the students create a concept map, using the following phrases and formulas. They can use those followed by the number 2 twice.

- Nonelectrolyte solution
- Vapor pressure lowering (2)
- $\Delta T_b = K_b\, m$
- $\Delta T_f = ik_f\, m$
- Freezing point depression (2)
- Osmotic pressure (2)
- $\pi = MRT$
- Electrolyte solution
- $\Delta T_b = ik_b\, m$
- $\Delta P = X_{\text{solute}}\, P°$
- Colligative properties
- $\Delta P = iX_{\text{solute}}\, P°$
- $\pi = iMRT$
- Boiling point elevation



1. Colligative properties
2. Nonelectrolyte solution
3. Electrolyte solution
4. Vapor pressure lowering
5. Freezing point depression
6. Boiling point depression
7. Osmotic pressure
8. Vapor pressure lowering
9. Freezing point depression
10. Boiling point depression

11. osmotic pressure
12. $\Delta P = X_{solute} P^\circ$
13. $\Delta T_f = k_f m$
14. $\Delta T_b = k_b m$
15. $\pi = MRT$
16. $\Delta P = iX_{solute} P^\circ$
17. $\Delta T_f = ik_f m$
18. $\Delta T_b = ik_b m$
19. $\pi = iMRT$

Assessment

Assess each student's work throughout section 1.9 (colligative properties of solution). To do this, observe and record how every student participates in discussing Activities 1.12 – 1.14 and in presenting the opinion of the group to the rest of the class. In addition, ask the students the meaning of colligative properties and have them state the colligative properties of solutions. Encourage them to do all the exercises and activities in the topic. Check their answers.

From your observations, make sure that students working at the minimum requirement level are able to: define colligative properties of solvents that are affected by the formation of a solution; state Raoult's law; explain the changes in the colligative properties of a solvent when a certain solute is added; calculate vapor pressure, boiling point, and the osmotic pressure of a solution; and compare and contrast changes in colligative properties of electrolytic and nonelectrolytic solutions.

Praise students working above the minimum requirement level. Assist students working below the minimum requirement level, either by arranging extra lesson time or giving them additional exercises.

Additional Questions

* 1. How many grams of sucrose ($C_{12}H_{22}O_{11}$) must be dissolved in 550 g of water to give a solution with a vapor pressure 3.0 mm Hg less than that of pure water at 20°C if the vapor pressure of water at 20°C is 17.5 mm Hg?

* 2. An ideal solution is prepared by mixing 30 g of methanol (CH_3OH) and 45 g of ethanol (CH_3CH_2OH) at 20°C. The vapor pressure of pure ethanol and pure methanol at this temperature are 44 mm Hg and 94 mm Hg, respectively.
   a) What is the partial vapor pressure of ethanol and methanol above the solution at 20°C?
   b) What is the total vapor pressure above the solution $P_{soln} = 68.5$ mm Hg

3. Which of the following two aqueous solutions has a) a higher boiling point, b) a higher freezing point c) a lower vapor pressure: 0.35 m $CaCl_2$ or 0.9 m urea?

4. Rank the following aqueous solutions, 0.10 M$NaNO_3$, 0.20m glucose and 0.10 m $CaCl_2$, in order of increasing
   a. Osmotic pressure b. Boiling point
   c. freezing point d. vapor pressure

* 5. The boiling point of ethanol (C_2H_5OH) is 78.5°C. What is the boiling point of a solution of 3.4 g vanillin (molar mass = 152 g/ mol) in 50.0g ethanol? (K_b of ethanol = 1.22°C/m)

6. Two beakers, one containing an 8.0M LiCl aqueous solution and the other containing pure water, are placed under a tightly sealed bell jar at room temperature. After a few months one beaker is completely dry, whereas the other has increased in volume.
   a. Is the volume of the solution or that of water increased?
   b. Which beaker is found completely dry? Why?

7. A solution of an unknown organic solid in 27.8 g of the solvent diphenyl gave a freezing point depression of 1.56°C. Analysis of the compound showed that it contained 40%C, 6.7% H and 53.3% O. What is
   a. the molar mass of the compound,
   b. the molecular formula of the compound? (K_f for the solvent diphenyl is 8.0°C/m.)

Answers to Additional Questions

1. 2.2×10^3g

2. a. $P_{Methanol} = 46$mmHg, $P_{Ethanol} = 22.5$ mmHg
   b. $P_{Soln} = 68.5$ mmHg

3. a. $CaCl_2$ solution b. Urea Solution c. $CaCl_2$ solution

4. a. $\pi_{NaNo_3} = \pi_{glucose} < \pi_{CaCl_2}$
   b. Boiling points of $NaNO_3$ and glucose are equal. The boiling point of $CaCl_2$ solution is higher than that of $NaNO_3$ and glucose solutions.
   c. The freezing point of $CaCl_2$ solution is lower than that of $NaNO_3$ and glucose. Solutions of $NaNO_3$ and glucose have the same freezing point.
   d. VP of $CaCl_2$ < VP of $NaNO_3$ = VP of glucose.

5. 79°C

6. a. The volume of the solution
   b. The beaker containing pure water. This is because of the diffusion of water from the beaker containing pure water to the beaker filled with LiCl solution.

7. a. 120 g b. $C_4H_8O_4$

Answers to Exercises

Exercise 1.14

1. Since solute particles take positions originally occupied by solvent particles, the number of solvent particles closer to the surface of the solution decreases. This condition, in turn, reduces the number of solvent particles entering into the vapor phase so that the vapor pressure of the solution becomes less than that of the pure solvent.
2. a. $X_{\text{ethyleneglycol}} = 0.290$ b. $X_{H_2O} = 0.710$
3. $\Delta P = 0.461$ torr

Exercise 1.15

1. b.p = 221°C
2. b.p = 102.8°C

Exercise 1.16

1. f. p = -3.72°C
2. molar mass = 255.4 g/mol

Exercise 1.17

1. molar mass = 4.20×10^4 g/mol
2. $\pi = 7.6$ atm.

Exercise 1.18

i = 3.5

Answers to Review Questions and Problems

Part I

1.	B	2.	C	3.	B	4.	C	5.	A
6.	B	7.	A	8.	D	9.	A	10.	C
11.	B	12.	B	13.	C	14.	C	15.	C

Part II

For questions 16 – 20 refer to students' textbook.

21. a. 0.513 M NaCl

    b. 2.03 M $LiClO_4.3H_2O$

22. 5.0×10^2 m, 18.3M

23. 375cm^3

24. a. A saturated solution is one that is in equilibrium with undissolved solute.

    b. A supersaturated solution can be prepared by saturating a solution at high temperature and carefully cooling it.

25. Refer to students' textbook

26. Refer to students' textbook

27. Refer to students' textbook

28. Ethanol: 30.0 mmHg

    1 - propanol: 26.3 mmHg

29. molar mass = 123 g/mol

30. $\pi = 4.40$ atm

31. Solvent or solution f.p $= 4.37^{o}$C

32. Refer to the student's textbook

33. Solvent or solution $K_f = 5.08^{o}$C kg/mol

34. $i = 2.75$

35. 5.4×10^5

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UNIT 2

ACID - BASE EQUILIBRIA

UNIT OVERVIEW Total of Periods allotted: 26 periods

In this unit, two important classes of compounds, acids and bases, are described. In section 2.1, the three acid-base concepts, Arrhenius, Brønsted-Lowry and Lewis concepts are given due emphasis. In section 2.2 the ionic equilibria of weak acids and bases will be studied. Under this section the concentration of the ionized molecule of a weak acid (HA), hydrogen ion, and conjugate base (A^-) of the weak acid are determined from Ka, the dissociation constant of weak acid.

The extent of the ionization of a weak acid in solution can be reduced by adding a compound containing the conjugate base of the weak acid (the common ion effect) or by adding another acid. This concept will be studied in section 2.3. This section also includes a buffer solution, a solution containing a mixture of an acid and its conjugate base, or of a base and its conjugate acid. Section 2.4 deals with hydrolysis of salts. The presence of a salt in water may affect the pH of a solution because the cations or the anions of the salt may react with water and produce an in balance of H_3O^+ and OH^- ions. The last section, section, 2.5, emphasizes acid-base indicators and acid-base titration.

In this unit, there are suggested activities, experiments and exercises. Due emphasis should be given to these activities in order to make the active learning method practical. To teach contents in this unit, question-and-answer, gapped lecture, group discussion and experiment are the suggested methodologies.

Unit Outcomes

After completing this unit, students will be able to:

- *understand the Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases;*
- *understand the dissociation of water, weak monoprotic and polyprotic acids, and bases;*
- *know how to solve equilibrium problems involving concentration of reactants and products,* K_a, K_b, pH and pOH;
- *understand the common-ion effect, buffer solutions, hydrolysis of salts, acid-base indicators and acid-base titrations;*

- *explain how buffering actions affect our daily lives, using some examples;*
- *determine the equivalents of acids and bases, respectively, that are required to neutralize specific amounts of base and acid;*
- *predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic or neutral;*
- *know how to solve problems involving concentration and pH of acid base titration;*
- *describe scientific enquiry skills encountered in this unit: classifying, communicating, asking questions, applying concepts and making generalizations.*

Main Contents

2.1 Acid - Base Concepts

2.2 Ionic Equilibria of Weak Acids and Bases

2.3 Common Ion Effect and Buffer Solution

2.4 Hydrolysis of Salts

2.5 Acid - Base Indicators

2.1 ACID - BASE CONCEPTS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *define acid by the Arrhenius concept;*
- *explain why a proton exists, bounded to water molecules, as H_3O^+, in all acid-base reactions;*
- *give examples of Arrhenius acids;*
- *define base using the Arrhenius concept;*
- *give examples of Arrhenius bases;*
- *define acid using the Brønsted-Lowry concept;*
- *give examples of Brønsted-Lowry acids;*
- *define base using the Brønsted-Lowry concept;*
- *give examples of Brønsted-Lowry bases;*

- *explain what conjugate acids and conjugate bases are;*
- *identify the acid-base conjugate pairs from a given reaction;*
- *write an equation for self-ionization of water and ammonia;*
- *explain what is meant by amphiprotic species;*
- *give examples of reactions of amphiprotic species;*
- *define an acid by using Lewis concept;*
- *give examples of Lewis acids;*
- *define a base, using the Lewis concept;*
- *give examples of Lewis bases.*

Forward Planning

Read the contents in section 2.1 (Acid-Base Concepts) to familiarize yourself with acid-base concepts. Plan which contents and activities you are going to deal with during each period so that are of the content in the section will be covered within five periods. In your plan, indicate the time you allot for group discussion, presentation, harmonizing concepts, stabilization and other activities you may perform in each period.

Also plan how to manage students during discussions. You need to read the teacher's guide on this section to get information about the methodologies to be implemented and to gain more ideas about Activities 2.1 – 2.4.

Subject Matter Presentation

Students have some background knowledge of this concept. The *start-up activity* is given to begin the unit with what students have learned in Grade 10. Harmonize the observation of the students by giving the generalization that acids change blue litmus to red, while bases change red litmus to blue.

The Arrhenius Concept of Acids and Bases

To teach this topic, group discussion and gapped lecture are the suggested methodologies.

As you present the lesson on *Acid - Base Concepts*, start with Activity 2.1. The activity enables students to recall the acid-base concepts they learned in grade 10. So, let the students discuss **Activity 2.1** in their groups for a few minutes by recalling what they have learned in Grade 10. In facilitating the lesson on *The Arrhenius Concept of Acids and Bases,* you may involve your students in discussion, using **Activity 2.2.** Let them discuss, in their groups, for a few minutes and then have students from different groups suggest their opinions to the class. Then, harmonize concepts suggested by the students with the facts. Before you go to the other concept, discuss the limitations of the *Arrhenius concept* with emphasis on the following points.

- ✓ The theory can be applied only to reactions that occur in water, because it defines acids and bases in terms of what happens when compounds dissolve in water.
- ✓ It does not give any reason why some compounds, such as HCl, dissolve in water to give acidic solutions, whereas others such as CH_4 do not.
- ✓ The theory states that only compounds that contain OH^- ions can be classified as Arrhenius bases. It cannot explain why other compounds, such as Na_2CO_3, have the characteristic properties of bases.

Brønsted-Lowry Concept of Acids and Bases

Use group discussion and gapped lecture to teach this topic.

As you proceed to the next concept, start with Activity 2.3. The activity helps students to recall the Brønsted - Lowry concept of acids and bases and discover if there is a difference between this concept and that of Arrhenius. Therefore, let the students do Activity 2.3 in groups for a few minutes to define the Brønsted-Lowry concept and compare it with that of Arrhenius definition. Encourage some groups to present their opinion to rest of the class. To harmonize their discussion, emphasize that the Brønsted - Lowry definition defines acids as proton donors and bases as proton acceptors. Illustrate this by taking a hypothetical acid, HA, which releases a proton and its conjugate base, A^- on dissociation.

$$\underset{\text{Acid}}{HA} \longrightarrow \underset{\text{Proton}}{H^+} + \underset{\text{Conjugate base}}{A^-}$$

After accepting a proton, a conjugate base reforms the acid. When a base, B, accepts a proton, it produces its conjugate acid, HB^+

$$\underset{\text{Base}}{B} + \underset{\text{Proton}}{H^+} \longrightarrow \underset{\text{Conjugate acid}}{HB^+}$$

If a conjugate acid releases a proton, it re-forms the base. In the course of your presentation, discuss that the Brønsted - Lowry approach can be applied to proton transfer reactions that do not happen in an aqueous medium, such as the reaction of ammonia gas with hydrogen chloride gas.

$$NH_3(g) + HCl(g) \longrightarrow NH_4^+ + Cl^-$$

In this reaction, the HCl molecule donates a proton to the ammonia molecule, making HCl a Brønsted - Lowry acid. The ammonia molecule accepts the proton, making it a Brønsted-Lowry base. A Brønsted - Lowry acid must have a proton to donate, and a Brønsted-Lowry base must have a lone pair of electrons in order to accept a proton.

Before you conclude your introduction of the Brønsted-Lowery concept of acids and bases, let the students get information about the points (2-4) in activity 2.3. Tell them that all Arrhenius acids, such as HCl, HBr, HI, $HClO_4$, are Brønsted-Lowery acids. But, the reverse is not true. This is because Brønsted-Lowry acids include all substances that can donate a proton (H^+) in aqueous solution or in any other medium or in the gaseous state. Inform them that conjugate acids formed after bases accept a proton are Brøonsted-Lowry acids such as H_3O^+, NH_4^+ etc. Let them also know that all Arrhenius bases commonly contain only one Brønsted-Lowry bases, which is the OH^-. NaOH, KOH etc. are not Brønsted-Lowry bases. Only the OH^- they release is a base. For example, $NH_2^-, S^{-2}, F^-, CN^-$ are Brønsted-Lowry bases.

Mention the inverse relationship between the strengths of acids and their conjugate bases. The conjugate bases of strong acids tend to be relatively weak bases and the conjugate bases of weak acids tend to be strong bases. Similarily, the stronger the base the weaker is its conjugate acid and the weaker the base, the stronger is its conjugate acid.

Autoionization of Substances and Amphiprotic Species

Notice that a chemical species is **amphiprotic** if it either donates or accepts a proton, and a chemical species is **amphoteric** if it reacts with both acids and bases. You may ask the students to give examples of amphoteric species from the knowledge they gained in lower grades.

Lewis Concept of Acids and Bases

It is advisable to use group discussion and gapped lectures for this section.

As you move to present the lesson on the *Lewis Concept* of *Acids and Bases*, encourage the students to do **Activity 2.4**. Start with Activity 2.4. The activity is aimed at enabling students to recall that the Lewis is a concept of acids and bases and to discover how it differs from the Brønsted-Lowry concept. Let group representatives explain the Lewis acid-base concept to the class. Harmonize their discussion by defining the Lewis acid and Lewis base as follows.

A Lewis acid is an electron-pair acceptor, and a Lewis base is an electron-pair donor. After a Lewis acid accepts an electron pair from a Lewis base, it forms an acid-base adduct. Lewis acids are electron-deficient species such as cations, molecules that have an atom with less than eight outermost electrons and molecules that can expand their octets. Lewis bases are electron-rich species such as anions and molecules with atoms with one or more lone pairs.

Let students also get the following information in relation to the points in activity 2.4. Brønsted-Lowry acids are proton (H^+) donors. The Lewis concept of acids considers H^+ as one kind of acid only. According to the Lewis concept, an acid is any species that has a vacant orbital or unfilled shell and can form a coordinate covalent (dative) bond by sharing an electron pair from any other species.

The Brønsted-Lowry concept considers species as bases only if they can accept a proton (H^+). For the base to accept a proton, it should have a lone pair of electrons. Thus, Brønsted-Lowry bases are also Lewis bases. However, Lewis bases include all species having a lone pair of electrons and are capable of forming a dative bond by supplying the lone pair not only with H^+, but also with other species. The Brønsted-Lowry concept restricted the definition of acids and bases to proton transfer only. It doesn't explain why species like CO_2, SO_2, Cu^{2+} etc. function as acids even if they do not contain a proton.

Let the students practice how to identify Lewis acids and Lewis bases in reactions similar to the ones given in the following example.

Example Identify the Lewis acid and Lewis base in each of the following reactions.

a) $BF_3 + PH_3 \longrightarrow F_3BPH_3$

b) $H^+ + H_2O \longrightarrow H_3O^+$

c) $Cu^{2+} + 4{:}NH_3 \longrightarrow \left[Cu({:}NH_3)_4\right]^{2+}$

Solutions:

a. PH_3 is the electron – pair donor, and therefore it is the Lewis base. BF_3 is the Electron-pair acceptor, and therefore it is the Lewis acid.

b. H_2O is the Lewis base, and H^+ is the Lewis acid.

c. NH_3 is the Lewis base, and Cu^{2+} is the Lewis acid.

Assessment

You can assess each student's work as you proceed through the acid-base concept (section 2.1). To do this, you need to observe and record everything about the performance of each student. You may make a record in your permanent performance list in relation to how each student participates in:

- discussing Activities 2.1 – 2.4
- presenting ideas or opinions of the group after discussion
- answering questions raised during gapped lectures
- attempting activities given in between gapped lectures

Give them Exercises 2.1 – 2.4 as class work or homework. Check their work and record their performance. You may also ask students oral questions during stabilizing concepts in each period and record the responses of the students.

From the cumulative record you have, make sure that the competencies suggested for section 2.1 are achieved by most of the students praise students working above the minimum requirement level and give them extra activities. With regard to students working below the minimum requirement level, you may arrange extra lesson time or give them additional exercises. You may use questions recommended to them from the Additional Questions in this section in the teacher's guide. If necessary, you can prepare additional questions of your own.

Additional Questions

*1. Identify the acid-base conjugate pairs in each of the following reactions.

a. $H_2PO_4^- + NH_3 \rightleftharpoons HPO_4^{2-} + NH_4^+$

b. $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$

c. $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

d. $HClO_4 + H_2SO_4 \rightleftharpoons ClO_4^- + H_3SO_4^+$

e. $HPO_4^{2-} + HSO_4^- \rightleftharpoons H_2PO_4^- + SO_4^{2-}$

f. $HSO_3^- + CH_3NH_2 \rightleftharpoons SO_3^{2} + CH_3NH_3^+$

*2. Write the formulas of the conjugate bases of

a. H_2S b. HCOOH c. HClO

d. H_3PO_4 e. NH_4^+

*3. Write the formulas of the conjugate acids of

a. HSO_3^- b. HCO_3^- c. CN^-

d. SO_4^{2-} e. OH^-

4. The acid strength of HCN < CH_3COOH < HF < HNO_2 < HCOOH. What is the correct increasing order in base strength of the conjugate bases of the acids?

5. Strong acids have weak conjugate bases and weak acids have strong conjugate bases. Why?

* 6. Identify the amphiprotic species among the following

a. NH_3 b. HCO_3^- c. HS^- d. HPO_4^-

e. H_2O f. NH_4^+ g. $H_2PO_4^-$

7. What is
   a. the strongest acid and
   b. the strongest base that can exist in aqueous solution?

*8. Indicate whether the following would be expected to serve as either a Lewis acid or a Lewis base:

a. $AlCl_3$ b. OH^- c. Br^- d. CO_2

e. Fe^{3+} f. Cu^{2+} g. PCl_5

Answers to Additional Questions

1. a. $H_2PO_4^- / HPO_4^{2-}$; NH_3 / NH_4^+
   b. CO_3^{2-} / HCO_3^-; H_2O / OH^-
   c. CH_3COO^- / CH_3COOH; H_2O / OH^-
   d. $HClO_4 / ClO_4^-$; $H_2SO_4 / H_3SO_4^+$
   e. $HPO_4^{2-} / H_2PO_4^-$; HSO_4^- / SO_4^{2-}
   f. HSO_3^- / SO_3^{2-}; $CH_3NH_2 / CH_3NH_3^+$
2. a. HS^- b. $HCOO^-$ c. ClO^-
   d. $H_2PO_4^-$ e. NH_3
3. a. H_2SO_3 b. H_2CO_3 c. HCN
   d. HSO_4^- e. H_2O
4. $CN^- > CH_3COO^- > F^- > NO_2^- > HCOO^-$
5. If the acid is strong, its conjugate base has little tendency to accept a proton and vice-versa.
6. All are amphiprotic, except d and f.
7. a. H_3O^+ b. OH^-
8. Lewis bases: b and c
   Lewis acids: a, d, e, f and g.

Answers to Exercise 2.1

a. Arrhenius acid

b. Cannot be classified as an Arrhenius acid or an Arrhenius base

c. Arrhenius base

d. Arrhenius base

Answers to Exercise 2.2

a. $NH_3 + HCO^-_3 \rightleftharpoons NH_4^+ + CO_3^{2-}$

Base Acid Conjugate acid Conjugate base

b. $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

Acid Base conjugate base Conjugate acid

c. $H_2O + SO_3^{2-} \rightleftharpoons OH^- + HSO^-_3$

Acid Base Conjugate base Conjugate acid

d. $CH_3COOH + H_2O \rightleftharpoons CH_3\ COO^- + H_3O^+$

Acid Base Conjugate base Conjugate acid

Answers to Exercise 2.3

1. a. Autoionization is a reaction that involves the transfer of a proton from one molecule to another molecule of the same substances.

   b. Amphiprotic species are molecules or ions that can act as a proton donor or proton acceptor.

2. a. HpO_4^{2-} b. H_2O

3. a. PO_4^{3-}, because it is conjugate base of a weak acid, HPO_4^{2-}.

   b. S^{2-}, because it is a conjugate base of a weak acid, HS^-.

   c. CO_3^{2-}, because it is a conjugate base of the weak acid HCO_3^-.

4. It restricted an acid-base concept only to proton transfer.

5. $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

   $NH_3 + NH_3 \rightleftharpoons NH_4^+ + NH_2^-$

Answers to Exercise 2.4

a. $SiCl_4 + 2Cl^- \rightarrow [SiCl_6]^{2-}$

Lewis acid Lewis base

b. $pF_5 + F^- \rightarrow [pF_6]^-$

Lewis acid Lewis base

c. $Cu^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4]^{2+}$

Lewis acid Lewis base

d. $CO_2 + H_2O \rightarrow H_2CO_3$

Lewis acid Lewis base

e. $Ni + 4CO \rightarrow Ni(CO)_4$

Lewis base Lewis acid

2.2 IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

Period Allotted: 9 periods

Competencies

At the end of this section, students will be able to:

- *describe the ionization of water;*
- *derive the expression of ion product for water,* K_w*;*
- *explain the effect of temperature on* K_w*;*
- *explain the effect of temperature on* K_w*;*
- *explain why water is a weak electrolyte;*
- *use* K_w *to calculate* $[H_3O^+]$ *or* $[OH^-]$ *in aqueous solution;*
- *define pH;*
- *define pOH;*
- *explain the relationship between pH and pOH;*
- *calculate pH from* $[H^+]$ *and* $[H^+]$ *from pH;*
- *calculate pOH from* $[OH^-]$ *and* $[OH^-]$ *from pOH;*
- *write an expression for the percent ionization of weak acids or weak bases;*
- *calculate the percent dissociation of weak acids and bases;*
- *write the expression for the acid-dissociation constant,* K_a*;*
- *calculate* K_a *for an acid from the concentration of a given solution and its pH;*
- *calculate* $[H^+]$ *and pH of an acidic solution from given values of* K_a *and the initial concentration of the solution;*
- *write the expression for the base-dissociation constant,* K_b*;*
- *calculate* K_b *for a base from the concentration of a basic solution and its pOH ;*
- *calculate the* $[OH^-]$ *and pOH of a basic solution from a given value of* K_b *and the initial concentration of the solution.*

Forward Planning

Read the contents given in the student's textbook on the ionic equilibria of weak acids and bases thoroughly and make the appropriate preparation. Design a plan that shows which contents and activities you are going to deal with in each period in such a manner that you can complete the entire contents in section 2.2 within nine periods.

In your plan, indicate the duration of time you may allot for group discussion, presentation, gapped lecture and other activities you perform in each period. In addition, read the teacher's guide on this section to help you get information about the methodologies you can implement to teach this section and to gain more ideas about Activities 2.5 – 2.8. Plan how to manage students during discussion and presentation.

Make sure that there is a pH meter or a universal indicator paper in your school chemistry laboratory. Prepare the materials required for students to do Activity 2.6. You may tell them to bring these materials before the period you intend to deal with this activity.

Subject Matter Presentation

Ionization of Water

Use group discussion and gapped lectures to teach the contents in this topic.

As you start the lesson on Ionization of Water, you may ask the in-text question. Continue your explanation by deriving the ion product of water, K_w, from self ionization of water. Then let the students know that, in any aqueous solution at 25°C, the product of $[H^+]$ and $[OH^-]$ must always equal 1.0×10^{-14}. Based on this expression, ask the students to write the relationship between $[H^+]$ and $[OH^-]$ in neutral, acidic and basic solutions.

Then proceed to Activity 2.5. The activity is designed to help students understand how they can write dissociation constant expressions for self-ionization of other substances and identify the strongest acid and base that can exist in liquid ammonia.

Encourage the students to derive the K_c for ammonia in analogous to that of K_w starting from self ionization of ammonia (**Activity 2.5**).

a. $2NH_3 \rightleftharpoons NH_4^+ + NH_2^-$

$$K_C = \frac{[NH_4^+]\,[NH_2^-]}{[NH_3^-]^2}$$

b. In liquid ammonia NH_4^+ is the strongest acid, and NH_2^- is the strongest base.

c. Here a solution with $[NH_4^+] > [NH_2^-]$ is acidic.

The pH Scale

It is preferable to implement gapped lecture to teach contents in this topic.

Start teaching this lesson by using Activity 2.6. The activity helps students to learn how to determine the pH of different substances using a pH meter. Let the students use a pH meter to do **Activity 2.6.** Tell them to collect the materials given in the table and determine whether they are acidic, basic or neutral from their pH values. Harmonize their discussion by comparing their results with the following table.

Sample	pH value	Acidic, basic or neutral
Beer	4–5	Acidic
Milk of magnesia	10.5	Basic
Tomato juice	4.0	Acidic
Lemon juice	2.4	Acidic
Drinking water	5.5 – 8.0	Maybe acidic, basic or neutral
Saliva	6.4 – 6.9	Slightly acidic

Then, continue with Activity 2.7, which is designed to help student realize how the relationship between pH, pOH and Pk_w is derived from the dissociation of water.

Let the students derive the expression: pH + pOH = p_{kw} = 14 at 25°C (**Activity 2.7**). This can be done from the relation

$$K_w = [H^+]\,[OH^-] = 1.0 \times 10^{-14} \text{ at } 25\,^{\circ}C$$

Take the negative logarithm of both sides

$$-\log k_w = -\log [H^+]\,[OH^-] = -\log (1.0 \times 10^{-14})$$

$$-\log k_w = -\log [H^+] + -\log [OH] = -\log (1.0 \times 10^{-14})$$

$$pKw = pH + pOH = 14$$

Using the examples discussed in the student's textbook or the additional examples given below, show the students how to calculate pH, pOH, $[H_3O^+]$, and $[OH^-]$.

Example What are the pH and pOH of a 0.025M HNO_3 solution?

Solution: Nitric acid (HNO_3) is a strong acid. For any strong monoprotic acid, the concentration of the hydrogen ion, H^+, is the same as the concentration of the acid itself. $[H^+] = 0.025M$

$$pH = -\log [H^+] = -\log (0.025) = 1.60$$

$$pH + pOH = 14.00$$

$$1.60 + pOH = 14.00$$

$$pOH = 14.00 - 1.60 = 12.40$$

Measures of Strength of Acids and Bases in Aqueous Solution

It is advisable to use group discussion, gapped lecture, and question-and-answer as methodologies to teach the contents in the topic.
After introducing *Measures of the Strengths of Acids and Bases in Aqueous Solution*, continue with Activity 2.8. The activity helps students to recall strong and weak acids, and strong and weak bases. So, let your students get involved in presenting their work on Activity 2.8. It is recommended that this activity be assigned ahead of time. In connection with to this activity, explain to them that the strength of acids and bases can be measured by the concentration of hydrogen ion and hydroxide ion, pH and pOH, percent ionization, and information from student's textbook.

Illustrate, with several numerical examples, how to calculate the dissociation constant from the concentration of the solution and vice versa. Table 2.2 in the student's textbook lists some weak acids and their k_a values.

Important guidelines are given in the student's text to help the students to solve weak-acid dissociation problems. Explain to the students that such problems can be treated in two ways. First, by the approximation method. If the approximation is not valid, we use the quadratic equation. You can use the examples given in the students' textbook to illustrate the calculations.

Assessment

Assess each student's work as you proceed throughout section 2.2, Ionic equilibria of Weak Acids and Bases. You can do this by recording how every student is doing, using your permanent performance list. Your records can be based on how every student:

- participates in discussing and performing Activities 2.5 – 2.8
- participates in presenting opinions and findings of the group after discussion
- in answering questions raised during the process of harmonizing concepts or during gapped lectures
- in performing activities given between the gapped lectures

Give them Exercise 2.5, 2.6 and 2.8 as homework. Check their work and record their performances. You may also give them a test on this section. Correct it and record their achievements. From the record you have, check whether or not the competencies suggested for this section are achieved by most of the students. Praise students working above the minimum requirement level and give them Exercise 2.7 as additional exercise. For students working below the minimum requirement level, you may give the following exercise.

1. What is the hydroxide ion concentration in a solution that has 0.10M H_3O+?

2. A student prepared a 0.01M solution of formic acid, $HCHO_2$, and measured its pH using a pH meter. The pH at 25°C was found to be 2.38.

   a. Calculate Ka value for formic acid at this temperature.

   b. What percent of the acid is ionized in this 0.10M solution?

You can also give them the questions recommended for slow learners from the Additional Questions given in this teacher's guide.

Additional Questions

*1. The value of K_{ω} at 40°C is 2.92×10^{-14}. What is the pH of water at this temperature?

* 2. 0.10 M solutions of acetic acid, chlorous acid, hydrofluoric acid and nitrous acid contain the following equilibrium concentrations.

Acid	[Acid]	[Conjugate Base]	$[H_3O^+]$
CH_3COOH	0.099M	0.0013M	0.0013M
$HClO_2$	0.072M	0.028M	0.028M
HF	0.092M	0.0081M	0.0081M
HNO_2	0.093M	0.0069	0.0069M

   a. Arrange the solutions of these acids in order of increasing acid strength.
   b. Which of these acids have the strongest conjugate base?

* 3. A 0.1M solution of nitrous acid (HNO_2) is 7.1% ionized at equilibrium. What is the
   a. $[H^+]$ and pH,
   b. Ka value of HNO_2
   c. $[OH^-]$ and pOH

4. What are $[H_3O^+]$, $[OH^-]$ and pOH in a solution of pH = 9.78?

5. The pH of a 0.012M solution of a weak based, BOH, was determined to be 11.40. What is the K_b value of the base?

* 6. The Ka values of CH_3COOH, HCOOH, HOCl and HOBr are 1.8×10^{-5}, 1.7×10^{-14}, 3.0×10^{-8} and 2.5×10^{-9}, respectively. A 0.06 M solution of which acid solution
   a. is the most acidic
   b. has the smallest pH value
   c. contains the lowest concentration of H^+
   d. the lowest pOH value?

Answers to Additional Questions

1. $P^H = 6.77$
2. a. $CH_3COOH < HNO_2 < HF < HClO_2$ b. CH_3COOH
3. a. $[H^+] = 7.1 \times 10^{-3}M$; $P^H = 2.15$ b. $Ka = 5.4 \times 10^{-4}$
   c. $[OH] = 1.4 \times 10^{-12}M$; $P^{OH} = 11.85$
4. $[H_3O^+]= 1.7 \times 10^{-10}M$; $[OH] = 6.0 \times 10^{-5}M$; $P^{OH} = 4.22$
5. $K_b = 6.6 \times 10^{-4}$
6. a. HCOOH b. HCOOH c. HOBr d. HOBr.

Answer to Exercise

1. $[OH^-] = 1.0 \times 10^{-13}$ M
2. a. $K_a = 1.8 \times 10^{-4}$ b. 4.2%

Answers to Exercise 2.5

1. a. $[H^+] = 1.0 \times 10^{-10}$, basic
   b. $[H^+] = 1.0 \times 10^{-6}$, acidic
   c. $[OH] = 1.0 \times 10^{-7}$, neutral
2. a. $5.0 \times 10^{-10}M$
   b. $5.0 \times 10^{-7}M$
   c. $1.0 \times 10^{-8}M$
3. 1.49×10^{-12}, basic
4. $[H^+] = [OH^-]$
5. Because it is dissociates only slightly.

Answers to Exercise 2.6

1. $[H^+] = 6.6 \times 10^{-10}$

   $[OH^-] = 1.5 \times 10^{-5}$

   pOH = 4.82

2. For 2.0M HNO_3

   $[H^+] = 2.0$ M, pH = -0.30, pOH = 14.30, $[OH^-] = 5.0 \times 10^{-5}$ M

   For 0.30 M HNO_3

   $[H^+] = 0.3$M, pH = 0.52, pOH = 13.48, $[OH^-] = 3.3 \times 10^{-14}$ M

   For 0.0063 M HNO_3

   $[H^+] = 0.0063$ M, pH = 2.20, pOH = 11.80, $[OH^-] = 1.6 \times 10^{-12}$ M

Answers to Exercise 2.7

a. $HNO_2 + H_2O \rightleftharpoons NO_2^-(aq) + H_3O+(aq)$

b. $X = [H_3O^+] = 4.0 \times 10^{-3}$ M

Test for approximation $\frac{4.0 \times 10^{-3}}{0.036} \times 100 = 11\%$

c. Using quadratic equation

$X = [H_3O^+] = 3.8 \times 10^{-3}$ M

d. pH = 2.42

Exercise 2.8

$NH_3(aq) + H_2O(l) \rightleftharpoons NH4^+(aq) + OH (aq)$

a. $[NH_4^+] = [OH^-] = 2.7 \times 10^{-3}M$

b. Approximation $\frac{2.7 \times 10^{-3}}{0.40M} M \times 100 = 0.68\%$ Approximation valid

c. pOH = 2.57

pH = 11.43

2.3 COMMON ION EFFECT AND BUFFER SOLUTION

Periods Allotted: 9 periods

Competencies

At the end of this section, students will be able to:

- *define the common-ion effect;*
- *explain the importance of the common-ion effect;*
- *define buffer solution;*
- *give some common examples of buffer systems;*
- *explain the action of buffer solutions and its importance in chemical processes;*
- *calculate the pH of a given buffer solution; and*
- *demonstrate the buffer action of CH_3COOH/CH_3COONa.*

Forward Planning

Read the contents in the students' text thoroughly on common ion effect and buffer solutions. This will help you to recall Le Chatelier's principles, since it is related to the common ion effect. You also get information about buffer solutions. Next, prepare a plan of your own that shows the contents, activities and experiments you are going to deal with during each period so that you can cover the entire contents of the section within nine periods. In your plan, indicate the duration of time required for group discussion, presentation, gapped lecture and other activities you perform during every period. Read the teacher's guide on this section to get more information about Activities 2.9 and 2.10 and the methodologies you use to teach the contents.

Prepare the chemicals, apparatus and other materials required to perform Experiment 2.1.

Teaching Aids

Refer to the text for the materials required to carry out Experiment 2.1.

Subject matter Presentation

The Common ion Effect

You can use group discussion and question-and-answer methods to teach contents in this part.

As you start teaching The Common ion Effect, let the students review the Lechateliers principle by doing Activity 2.9. This activity helps the students to see the effect of concentration, temperature and pressure on chemical equilibrium. Let them read in the Grade 11 chemistry book and other references to do this activity. After they discuss and present their report to the class, harmonize their discussion.

Relate the common ion effect with Activity 2.9 by using the example given in the students' textbook. In this example, the addition of sodium acetate, $NaC_2H_3O_2$, to a solution of acetic acid, $HC_2H_3O_2$, at equilibrium causes the equilibrium to shift to the left, thereby decreasing the equilibrium concentration of $H^+(aq)$. This indicates that the dissociation of a weak electrolyte is decreased by adding to the solution a strong electrolyte that has an ion in common with the weak electrolyte.

Give them additional dissociation reactions of weak electrolytes and let them discuss the effect of adding strong electrolytes having common ions with the weak electrolytes. You can use question number 1 from the additional questions and let students suggest the effect of adding a solution of HCOOK to that of HCOOH. Then, harmonize their suggestion based on the answer given in this guide.

Buffer Solutions

Use group discussion, question-and-answer, and experiment as methodologies to teach this part.

Start you discussion on *Buffer Solution* by defining the term “buffer”. The acetic acid/sodium acetate solution can be use to explain the common ion effect. A buffer solution is a solution that changes pH only slightly when a small amount of a strong acid or a strong base is added.

Buffered solutions contain a conjugate weak acid-base pair. The acidic member consumes a small amount of added base, while the basic member consumes small amounts of added acid. In the students’ textbook, you can get information about how buffers work. Read this part and explain it to the students.

Then, continue with Activity 2.10. This activity helps students to know about the presence of buffer solutions in nature, such as in human blood.

Let the students do Activity 2.10 in groups and one or two groups present their findings to the rest of the class. After their presentation, harmonize their discussion by referring the following explanation. Explain briefly the use of buffered solutions by taking human blood as an example. Buffers are important in our blood and other body fluids. Cells in living organisms must maintain a proper pH in order to carry out essential life processes, primarily because enzyme function is sharply dependent on pH. The normal pH value of blood plasma is 7.4. Sustained variations of a few tenths of pH unit can cause severe illness or death. Decrease in blood pH causes the condition called acidosis which can be brought on by heart failure, kidney failure, diabetes, persistent diarrhea or other factors. Similarly, alkalosis, characterized by an increase in the pH of blood may result from severe vomiting, excessive breathing (hyperventilation), and others. The major buffer system used to control the pH of blood is the carbonic acid-bicarbonate buffer system. Carbonic acid, H_2CO_3, and bicarbonate, HCO_3^-, form a conjugate acid base pair. In addition, carbonic acid can decompose into carbon dioxide gas and water.

Help the students to demonstrate experiment 2.1. From the experiment, students can observe that the addition of HCl affects the unbuffered solution than the buffered solution.

Illustrate how to calculate the pH of buffer solutions by using the example given in the textbook and the following additional examples.

Note for the Teacher

The *Henderson-Hasselbalch equation* may be applied to solve buffer-system problems. This equation can be derived by considering the acid dissociation reaction of HA:

$$HA \rightleftharpoons H_3O^+ + A^-$$

Rearranging this expression yields:

$$\left[H_3O^+\right] = K_a \frac{[HA]}{\left[A^-\right]} \Rightarrow K_a = \frac{\left[H_3O^+\right]\left[A^-\right]}{[HA]}$$

Taking the logarithm of both sides of the above equation gives:

$$\left[H_3O^+\right] = K_a \frac{[HA]}{\left[A^-\right]}$$

$$\log\left[H_3O^+\right] = \log K_a + \log\frac{[HA]}{\left[A^-\right]}$$

$$-\log\left[H_3O^+\right] = -\log\ Ka + -\log\frac{[HA]}{\left[A^-\right]}$$

$$pH = pK_a - \log \frac{[HA]}{\left[A^-\right]}$$

$$pH = pK_a + \log \frac{[HA]}{\left[A^-\right]}$$

This equation is called the Henderson-Hasselbalch equation.

Example

a. What is the pH of a buffer solution that is 0.25 M in HF and 0.50 M in NaF?
b. What is the final pH if 3.0 mL of 2.0 M HCl is added to 100 mL of the buffer solution in (a) above?

Solution:

a. To determine the pH in part (a), we will use the same method that we used in introducing the common ion effect, although we could use the Henderson-Hasselblach equation.

Because NaF is a strong electrolyte and dissociates completely, 0.50 M NaF gives 0.50 M F^-. [HF] = 0.25 M. Substituting these values in a Henderson - Hasselbalch equation.

$$pH = pK_a + \log\frac{[\text{Conjugate base}]}{[\text{Acid}]}$$

$$pH = -\log 6.8 \times 10^{-4} + \log \left(\frac{0.50}{0.25}\right)$$

$$= 3.15 + \log \left(\frac{0.50}{0.25}\right) = 3.45$$

b. Since HCl is a strong electrolyte, it is completely ionized to give H^+, which consumes F^- ions and generates additional HF. The number of moles of H^+ added in the form of HCl is:
$0.003\ L \times 2.0\ M = 0.006\ mol\ H^+$.

The number of moles of HF originally present in 100 mL of buffer is:

$0.100\ L \times 0.25\ mol/L = 0.025\ mol\ HF$

The number of moles of F^- originally present in 100 mL of buffer is:

$0.100\ L \times 0.50\ mol/L = 0.050\ mol\ F^-$

After the addition of 0.006 moles of H^+, 0.006 mol H^+ consumes 0.006 mol F^- and

generates 0.006 mol HF. Therefore, the number of moles of HF is:

$0.025 + 0.006 = 0.031$ mol.

The number of moles of F^- is:

$0.050 - 0.006 = 0.044$ mol.

Then we find the [HF] and $[F^-]$

in the total volume of 3.0 mL + 100 mL = 103 mL = 0.103 L.

$$[HF] = \frac{0.031\ mol}{0.103\ L}$$

$$[F^-] = \frac{0.044\ mol}{0.103\ L}$$

Using the equation

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$= 3.15 + \log\left(\frac{0.044\ mol/L \big/ 0.103\ L}{0.031\ mol \big/ 0.103\ L}\right) = 3.30$$

Assessment

Assess each student's work as you proceed through the contents of section 2.3. Your close follow-up is crucial. See how every student is working during the teaching-learning process. To do this, use your permanent performance list to record how every student.

- Participates in discussing Activities 2.9 and 2.10

- participates in presenting opinions of the group to the rest of the class
- answers questions raised during the process of harmonizing concepts, mini-lectures and stabilization
- takes part in performing Experiment 2.1
- participates in presenting results and observations from the experiment

Give Exercise 2.9 to slow learning students and Exercise 2.10 to gifted students. Check their exercise books and rescore their results.

From the record you have, check whether or not the competencies suggested for section 2.3 are accomplished by most of the students. Praise students working above the minimum requirement level and encourage them to continue working hard. Arrange extra lesson time or give additional exercises to students working below the minimum requirement level. You may use questions recommended to them from the Additional Questions given in the teacher's guide to help these students catch up with the rest of the class.

Additional Questions

* 1. What is the effect of adding potassium formate, HCOOK, to a 0.1M solution of formic acid, HCOOH? Explain the effect on
   a. the dissociation of HCOOH,
   b. the $[H^+]$,
   c. the pH value of the solution

* 2. What is the pH of a solution that consists of 0.55M HCOOH and 0.63M HCOONa? (Ka for HCOOH = 1.7×10^{-4})

* 3. What are the components of a buffer solution?

4. A buffer solution consists of 0.5M CH_3COOH and 0.5M CH_3COONa. Using a chemical equation, show how this buffer resists the change in pH upon addition of little amount of

   a. NaOH to the buffer b. HCl to the buffer

* 5. Which of the following pairs cannot form a buffer system?

   a. NaH_2PO_4/ Na_2HPO_4 b. H_2CO_3/ $NaHCO_3$

   c. HCOOH/HCOONa d. HCl/NaCl

   e. HCN/ NaCN f. H_2SO_4/$NaHSO_4$

   g. HOCl/NaOCl

6. Two buffer solutions, A and B, have the following concentrations of their components

Buffer	[CH_3COOH]	[CH_3COO^-]
A	0.5M	0.5M
B	0.1M	0.1M

a. What is the pH of both buffer solutions before adding any acid or base?

b. Which buffer resists the change in pH better upon the addition of 0.015 mole of solid NaOH to 100mL of each buffer solution? (Ka for $CH_3COOH = 1.8 \times 10^{-5}$)

Answers to Additional Questions

1. a. The dissociation HCOOH decreases due to the common ion $HCOO^-$, which shifts the equilibrium to the left.
   b. The [H^+] decreases. That is because it combines with $HCOO^-$ to form HCOOH.
   c. The P^H value of the solution increases.
2. $P^H = 3.8$
3. A weak acid and its conjugate base or a weak base and its conjugate acid.
4. a. $CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$
   b. $CH_3COO + H^+ \rightleftharpoons CH_3COOH$
5. ‘d’ and ‘f’ cannot form a buffer system.
6. a. $P^H = 4.745$ for both solutions.
   b. Buffer solution A, because the concentrations of its components are higher than those of B.

Answers to Exercises 2.9

a. pH = 4.92

b. pH = 2.72

Answers to Exercises 2.10

a. 4.74 b. 4.74 c. 4.75 d. 4.74

e. i. 10.99

ii. 3.0

2.4 HYDROLYSIS OF SALTS

Period Allotted: 2 periods

Competencies

At the end of this section, students will be able to:

- *define hydrolysis;*
- *explain why a salt of weak acid and strong base gives a basic solution;*
- *explain why a salt of strong acid and weak base gives an acidic solution;*
- *explain why salts of weak acids and weak bases give acidic, basic or neutral solutions.*

Forward Planning

Make the necessary preparations for the hydrolysis reaction of different types of salts with water. You may read the students' textbook and other relevant reference materials to do this. Read the teacher's guide to get information about activities 2.11 and 2.12 and to learn how you can present the contents.

Subject Matter Presentation

Use group discussion and gapped lectures to teach the contents in this section.

Start this subtopic with the definition of salt hydrolysis. We can assume that, when salts dissolve in water, they are completely ionized; nearly all salts are strong electrolytes. Consequently, the acid-base properties of salt solutions are due to the behavior of their cations and anions. Many ions are able to react with water to generate $H^+(aq)$ or $OH^-(aq)$. This type of reaction is often called **hydrolysis**.

After introducing these concepts, write the formula of potassium bromide, KBr on the board. Then ask the students to write the ionization of KBr in solution and predict what would the nature of its aqueous solution be. Invite one or two students to present their opinion. Following their presentation, tell them that KBr is a salt of strong acid (KBr), and strong base, KOH which ionizes as follows:

$$KBr(aq) \rightarrow K^+(aq) + Br^-(aq)$$

Neither K^+ nor Br^- reach with water. Thus a solution of KBr in water is neutral.

After that, continue with hydrolysis of salts of weak acids and strong bases. Start the lesson on this topic with Activity 2.11. The activity may help the students to identify the nature of the solution that a given salt forms by knowing the acid and base from which the salt is formed. Have the students discuss Activity 2.11 for a few minutes and some groups present their opinions to the class. After the presentations, harmonize concepts. Tell them that Na_2CO_3 is a salt of a strong base, sodium hydroxide, and a weak acid, carbonic acid. So, its solution in water will be basic. This is because the anions of the salt react with water and produce OH^- as follows:

$$CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$$

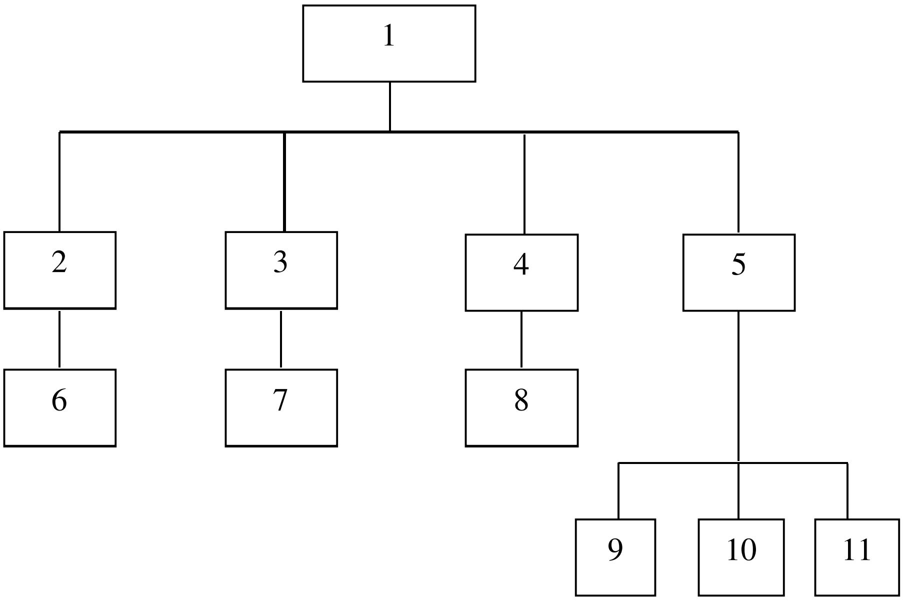
In relation to hydrolysis of salts of strong acids and weak bases, write the formula $(NH_4)_2SO_4$ on the blackboard and let students identify the parent acid and base, the ion that can be hydrolyzed and the nature of the solution. Invite one or two students to suggest their opinions. After their responses, have them know that $(NH_4)_2SO_4$ is a salt of a weak base, NH_3 and a strong acid, H_2SO_4. So it ionizes in aqueous solution into NH_4^+ and SO_4^{2-}. The NH_4^+ is a conjugate acid of a weak base and hydrolyzes to a greater extent and make the solution acidic. The hydrolysis reaction is:

$$NH_4^+ (aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

Next, continue with hydrolysis of salts of weak acids and weak bases. Start this lesson with Activity 2.12. The activity is given to students so that they can compare the K_a and K_b values of the anions and cations in order to know whether the salts under consideration form acidic, basic or neutral solutions. Have the students discuss this activity in groups for a few minutes and some groups present their opinion to the class. To harmonize concepts, tell them that, if $K_a > K_b$, the salt solution is acidic, and if $K_b > K_a$, the salt solution is basic, and if $K_a = K_b$ the salt solution is neutral. So, the solution of NH_4F is acidic because $K_a > K_b$, the solution of NH_4CNS is basic because $K_b > K_a$, and the solution of CH_3COONH_4 is neutral because $K_a = K_b$.

As a summary of the hydrolysis of salts, let students draw a concept map, using the following words and phrases. They can use those followed by number 2 in brackets twice. They should relate the type of salt and nature of solution the salt will form.

- Salt of strong acid and strong base
- Hydrolysis
- Neutral solution (2)
- Salt of strong acid and weak base
- Acidic solution (2)
- Salt of weak acid and strong base
- Basic solution (2)
- Salt of weak acid and weak base



1. Hydrolysis
2. Salt of strong acid and strong base
3. Salt of strong acid and strong base
4. Salt of weak acid and strong base
5. Salt of weak acid and weak base
6. Neutral solution
7. Acidic solution
8. Basic solution
9. Neutral
10. Basic
11. Acidic

Assessment

Use activity 2.11 and activity 2.12 as a means of assessment. Check whether they did these activities and record their results. You may give them the Additional Questions from the teacher's guide and check their work and record their achievements. Based on this, see whether or not the competencies suggested for section 2.4 are achieved by most of the students.

Additional Questions

* 1. What is salt hydrolysis?

* 2. Identify the salts whose aqueous solutions are acidic, basic or neutral among the following.

a. K_2SO_4 b. K_2CO_3 c. $(NH_4)_2SO_4$

d. NH_4Cl e. $NaNO_3$ f. NaCN

Answers to Additional Questions

1. The reaction of cation or anion or both with water.
2. 'a' and 'e' form neutral solution. 'b' and 'f' form basic solution. 'c' and 'd' acidic solutions.

2.5 ACID - BASE INDICATORS AND TITRATIONS

Period Allotted: 6 periods

Competencies

At the end of this section, students will be able to:

- *define acid-base indicators;*
- *write some examples of acid-base indicators;*
- *suggest a suitable indicator for a given acid-base titration;*
- *explain the equivalents of acids and bases;*
- *calculate the normality of a given acidic or basic solution;*
- *define acid-base titration;*
- *define end point;*
- *define equivalence point;*
- *distinguish between end point and equivalent point;*
- *discuss titration curves.*

Forward Planning

Read the student's textbook and other relevant references in order to understand acid-base indicators, equivalents of acids and bases and acid-base titration.

Make a plan that shows which contents, activities and experiment you may deal with during each period in such a manner as to complete the entire contents of this section within six periods. In your plan, indicate the duration of time allotted for group discussion, presentation, harmonizing concepts and other activities you perform in each period.

Prepare the chemicals, apparatus and other materials required to perform Experiment 2.2.

Teaching Aids

Refer to the student's text for the chemicals, apparatus and other materials required to perform experiment 2.2.

Subject Matter Presentation

Acid-Base Indicators

Use gapped lectures as the teaching method for this topic.

Here you may think of some relevant issue or question to brainstorm about with your students. Then continue by explaining to the students that an *acid - base indicator* is a weak acid or weak base with its corresponding conjugate pair. Mention that the molecular form of the indicator, HIn, has one color and that the ionic form, In^-, has a different color. The color that is observed in a solution of the indicator is controlled by the ratio of $[HIn]$ to $[In^-]$, which is determined, in turn, by the H^+ concentration in the solution.

Equivalence of Acids and Bases

It is preferable to use group discussion and question-and-answer as the methodologies to teach this topic.
Start teaching the lesson on this topic by using Activity 2.13. The activity is aimed at assisting students to realize what an equivalent of an acid and a base is and how to obtain their equivalent masses.

Let the students do Activity 2.13 in their groups for a few minutes and present their opinions to the class. After the presentations, harmonize concepts. An equivalent of an acid is the mass of the acid that releases one mole of H^+ in solution or reacts with one mole of OH^- completely. Likewise, an equivalent of a base is the mass of the base that produces one mole of OH^- or accepts one mole of H^+. For an oxidation-reduction reaction, one equivalent is the quantity of a substance that will react with or yield 1 mol of electrons. Note that the equivalent is defined in terms of a reaction, not merely in terms of the formula of a compound.

Discuss the difference between equivalent mass and molecular mass of a substance. The equivalent mass of a substance is the mass in grams of 1 equivalent of the substances. Illustrate this by using the following examples.

Sulphuric acid dissociates in water as follows:

$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

From the equation, it can be noted that one mole of H_2SO_4 releases two mole of H^+.
Thus, one mole of H_2SO_4 is 98 g/mol while its equivalent mass is 49 g/equivalent. From this, we can conclude that

$$\text{equivalent mass(g)} = \frac{\text{molar mass (g)}}{\text{number of equivalents per mole}}$$

Write the following chemical equations on the board and let the students determine the equivalent masses of $Ba(OH)_2$ and Fe_2O_3.

a. $Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$

b. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(\ell) + 3CO_2(g)$

Let the students know the difference between molarity and normality in relation to the difference between equivalent mass and molecular mass.

$$\text{Normality} = \text{Molarity} \times \text{Number of equivalents}$$

Let the students also know that during an acid-base neutralization reaction, one equivalent of an acid can neutralize one equivalent of the base. When neutralization is complete, the number of equivalents of the acid is exactly equal to that of the base. The formula; $N_1V_1 = N_2V_2$ proves this fact. They should also know that one mole of an acid may not necessarily neutralize one mole of a base.

Strong acid-Strong base Titration

It is advisable to use experiment, question-and-answer and group discussion to teach contents in this section.

Explain that *titration* is an analytical procedure in which a solution, generally of known concentration, is added gradually from a burette to another solution where the solute reacts. This is continued until the completion of the reaction, as indicated by an indicator which was added at the beginning of the titration.

Let the students perform Experiment 2.2 in their groups to learn what titration is and the procedures followed. Each group should write a laboratory report and present it to the class, describing their observations. After the presentations, harmonize concepts. In relation to the Observation and Analysis part of the experiment, you can tell them that:

1. Color change at the endpoint is from colorless to pink.
2. They can determine the volume of sodium hydroxide from the reading on the burette. Volume added = $v_i - v_f$, where v_i is initial volume and v_f final volume.
3. The normality can be calculated by using the relation: $N_1V_1 = N_2 V_2$.
4. Equivalence point is the point at which equivalent quantities of an acid and base are reacted. End point is the point at which the color change for the indicator occurs in a titration. Theoretically, in a given acid-base titration, the equivalence point and end point should coincide. However, there can be slight errors depending on the indicator used. For example, the pH range of color change for phenolphthalein is 8.2 – 10. In the titration of hydrochloric acid against NaOH solution, the equivalence point is at pH = 7.

Note for the Teacher

Illustrate the *titration of strong acid with strong base and weak acid with strong base* through the example given in the students' textbook. You may use the following additional explanation.

Inform the students that *equivalence point* is that point on the graph at which the number of moles of added OH^- equals the number of moles of H^+ originally present, whereas the end point is the point at which the color change of the indicator occurs in a titration. For a strong acid-base titration, the pH at the equivalent point is 7.00 because neither of the resulting ions reacts with water.

Weak Acid-Strong Base Titrations

Use the question-and-answer method to teach this topic. In the titration of *a weak acid with a strong base*, the resulting conjugate base of the weak acid is a strong base, and it will hydrolyze, yielding a basic solution at the equivalence point. Inform the students that in acid-base titration, the titration curve is a graph of the pH of the reaction solution versus the amount of titrant added.

Weak Base-Strong Acid Titrations

Start teaching the lesson by using Activity 2.14. The activity is designed to help students to practice drawing a titration curve for a weak base-strong acid titration.

Encourage the students to do **Activity 2.14**. Let them calculate the pH corresponding to the volume of HCl added and tabulate the pH values. From the pH values calculated, let them draw the titration curve pH-against-volume added. In relation to Activity 2.14, invite one student from the class to calculate the pH and draw the titration curve on the blackboard. After the student’s attempt, show them how to calculate the pH corresponding to the volume of HCl added, as shown in the following example.

Example

Before the addition of HCl:

Mole of NH_3 = volume of NH_3 × conc. of NH_3

$= 0.02\text{L}\ \ 0.100\text{ M}$

$= 0.002\text{ mole}$

$$[NH_3] = \frac{0.002\text{ mole}}{0.02\text{ L}} = 0.10\text{ M}$$

$$NH_3(aq) + H_2O\ (l) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

Initial, M	0.10 M	0	0
Charge, M	– x	+ x	+ x
Equilibrium, M	(0.10 M – x)	x	x

$$K_b = \frac{[NH_4^+]\ [OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.10\ -\ x}$$

Using approximation $0.10 – x \approx 0.10$

$$1.8 \times 10^{-5} = \frac{x^2}{0.10\ -\ x}$$

$$x^2 = 1.8 \times 10^{-5} \times 0.10$$

$$= 0.18 \times 10^{-5}$$

$$x = \sqrt{0.18\ \times\ 10^{-5}} = \sqrt{1.8\ \times\ 10^{-6}}$$

$$= 1.34 \times 10^{-3}$$

$$x = [OH^-] = 1.34 \times 10^{-3}$$

$$pOH = –\log\ (1.34 \times 10^{-3}) = 2.87,\ pH = 11.13$$

Example 2

When 25 mL of 0.10 M HCl:

This is the equivalence point. To calculate the pH at the equivalence, point:

Mole of HCl added

0.025L × 0.10 M = 0.0025 mol

Mol of NH_3 0.025 L × 0.10 M = 0.0025 mol

0.0025 mol HCl neutralizes 0.0025 mol NH_3, forming 0.0025 mol NH_4^+.

The total volume is 25 mL + 25 mL = 50 mL

$$[NH_4^+] = \frac{0.0025\ \text{mol}}{0.050\ \text{L}} = 0.05\ \text{M}$$

$$NH_4^+ + H_2O \rightleftharpoons NH_3^-(aq) + H_3O^+(aq)$$

Initial, M	0.05	0.00	0.00
Charge, M	– x	+ x	+ x
Equilibrium, M	0.05 – x	x	x

$$K_a = \frac{[NH_4^+]\,[H_3O^-]}{[NH_3]}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.05 - x}$$

Using approximation $0.05 - x \approx 0.05$

$$5.6 \times 10^{-10} = \frac{x^2}{0.05}$$

$$x^2 = 0.05 \times 5.6 \times 10^{-10}$$

$$x^2 = 2.8 \times 10^{-11}$$

$$x = \sqrt{28 \times 10^{-12}}$$

$$= 5.29 \times 10^{-6}$$

$$x = [H_3O^+] = 5.29 \times 10^{-6}$$

$$pH = -\log(5.29 \times 10^{-6})$$

$$= 5.28$$

Example 3

When 26 mL of 0.10 M HCl is added:

Mole of HCl added 0.026L × 0.10 M = 0.0026

Mole of NH_3

0.025L × 0.10 M = 0.0025 mol

Mole of HCl remained after neutralization (0. 0026 – 0.0025) mol = 0.0001 mol

Total volume 25 mL + 26 mL = 51 mL

Mole of H_3O^+ = mol of HCl remained

After neutralization = 0.0001 mol

$$[H_3O^+] = \frac{0.0001}{0.051} = 1.97 \times 10^{-3}$$

$$pH = -\log(1.97 \times 10^{-3})$$

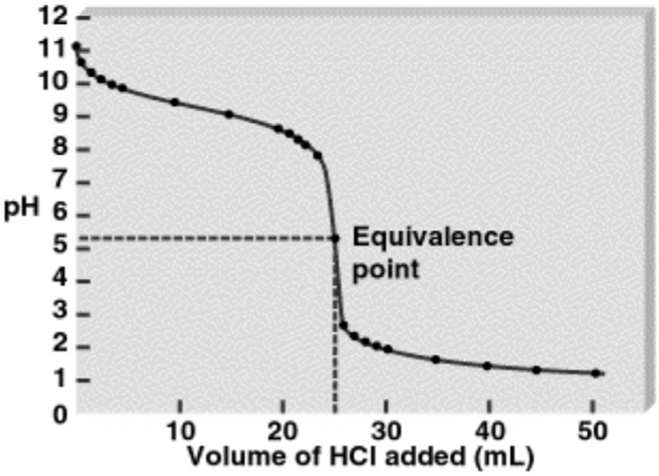
$$= 2.70$$

Let the students calculate the pH values corresponding to the volume of HCl added, according to the given example.

Compare their results with the following table.

Volume of HCl added (mL)	0.00	5.00	10.00	20.00	22.00	24.00	25.00	26.00
pH	11.13	9.86	9.44	8.66	8.39	7.88	5.28	2.70

Make sure that the curve drawn by students is similar to the one shown below and that the best indicator they suggested for the titration is methyl red.



Assessment

Assess each student's work throughout the section. Create a record of the performance of each student based on his/her participation

- in discussing Activities 2.13 – 2.14
- in presenting concepts after discussion
- in performing experiment 2.2
- in presenting the observation and result of the group to the rest of the class.

In addition to this, you can give them Exercises 2.11 – 2.13 as class work or homework. Check how the students attempted the exercises and record their achievements. From your records, make sure that the suggested competencies are accomplished.

Additional Questions

* 1. An indicator, HIn, has an ionization constant, K_{In} equal to 1×10^{-5}. If the molecular form of the indicator, HIn, is yellow, and the In^- is green, what is the color of a solution containing this indicator when its pH is

a. 3 b. 7 c. 5

2. What would the equivalent mass of phosphoric acid (H_3PO_4) be if it reacts with potassium hydroxide, KOH, according to the following equations?

a. $H_2PO_4 + KOH \rightarrow KH_2PO_4 + H_2O$,

b. $H_3PO_4 + 3KOH \rightarrow K_3PO_4 + 3H_2O$? (The molar of H_3PO_4 = 98 g.)

* 3. What volume of KOH solution is required to completely neutralize 50 mL of 0.5M H_2SO_4 solution?

4. In an acid-base, titration why is the pH at the equivalence point as follows:
   a. equals 7 for strong acid-strong base titration.
   b. is less than7 for strong acid-weak base titration.
   c. is greater than 7 for weak acid-strong base titration.

* 5. 50 mL of a 0.1M HNO_3 is titrated with a 0.05M KOH solution. What is the pH of the following:
   a. the acid solution before adding KOH solution.
   b. the acid solution after 50 mL of KOH solution has been added to it.

Answers to Additional Questions

1. a. yellow b. green c. yellowish-green

2. a. 98 g/eq b. 32.67 g/eq

3. 250 mL KOH solution

4. a. Neither the anion of the acid nor the cation of the base react with wale.

b. Because the cation divided from the base undergoes hydrolysis and increases $[H^+]$.

c. Because the anion derived from the acid undergoes reaction with water and increases the $[OH^-]$.

5. a. pH = 1 b. pH = 1.6

Answers to Exercise

Exercise 2.1

1. 75 mL of HNO_3
2. 8.0 N

Exercise 2.12

a. pH = 3 b. pH = 4 c. pH = 7 d. pH = 10

Exercise 2.13

a. pH = 4.57 b. pH = 8.72 c. pH = 12.22

Answer to Review Questions and Problems

Part I

1	D	2.	D	3.	C	4.	C	5.	C
6.	A	7.	C	8.	A	9.	B	10	B
11.	D	12.	C	13.	C	14.	B	15.	D

Part II

16. a. Arrhenius concept

An acid is the substance that releases H^+ or H_3O^+ in aqueous solution.

A base is the substance that releases the hydroxide ion, OH^-, in aqueous solution.

b. Brønsted-Lowry concept

An acid is a proton donor

A base is a proton acceptor

c. Lewis concept

An acid is electron pair acceptor

A base is electron pair donor

17. a. acid b. base c. base d. acid e. acid

18. a. H_2O b. HCl c. HOCl d. HCN
e. H_2CO_3 f. $H_2PO_4^-$ g. H_2S

19. a. HS^- b. $HCOO^-$ c. SO_3^{2-} d. SO_4^{2-} e. S^{2-}
f. NO_2^- g. CN^-

20. $H_2SO_4(aq) + H_2O(aq) \rightleftharpoons HSO_4^-(aq) + H_3O(aq)$

acid base 2 conjugate base conjugate acid

21. a. $H_2S(aq) + NH_3(aq) \rightleftharpoons HS^-(aq) + NH_4^+(aq)$

acid 1 base 2 base 1 acid 2

b. $CN^-(aq) + H_2O(l) \rightleftharpoons HCN(aq) + OH^-(aq)$

base 1 acid 2 acid 1 base 2

c. $H_2C_2O_4(aq) + H_2O(l) \rightleftharpoons HC_2O_4^-(aq) + H_3O^+ (aq)$

acid 1 base 2 base 1 acid 2

d. $HCO_3^- + OH^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2O(l)$

acid 1 base 2 base 1 acid 2

22. a. A substance is amphiprotic if it can act as either an acid or a base.

b. $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

23. b, c and f

24. a. $Cl^- < H_2O < OH^- < H^-$

b. $ClO_4^- < ClO_3^- < ClO_2^- < ClO^-$

c. $HTe^- < HS^- < PH_2^- < NH_2^-$

d. $ClO_2^- < BrO_2^- < IO_2^-$

25. b and d

26. a. $[OH^-] = 2.0 \times 10^{-12}$ M b. $[OH^-] = 7.7 \times 10^{-6}$ M

c. $[OH^-] = 1.0 \times 10^{-6}$ M

27. a. pH = 2.44 b. pH = 1.33

c. pH = 8.67 d. pH = 12.83

28. a. 2 b. 2.17 c. 1.98 d. 10.55

29. % dissociation = 5.67×10^{-3}, pH = 5.07, $[OH^-] = 1.18 \times 10^{-9}$ M

30. $[OH^-] = 4.78 \times 10^{-4}$ M, % dissociation = 47.8

31. $K_a = 1.5 \times 10^{-3}$

32. Refer to the students' text

a. page 88 b. salt containing Cl^- in excess c. page 88

33. a. Increase b. decrease

c. no change d. decrease

34. Because a buffer should contain either a weak acid and its conjugate base or a weak base and its conjugate acid.

35. pH = 8.88

36. $\left[CH_3COO^-\right]/\left[CH_3COO\right] = 0.58$

37. a. pH = 4.60

b. $Na^+(aq) + C_2H_3O_2^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow HC_2H_3O_2(aq) + Na^+(aq) + Cl^-(aq)$

c. $HC_2H_3O_2(aq) + Na^+(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l) + Na^+(aq)$

38. a. pH = 4.78 b. pH = 4.73 c. pH = 4.83

39. a The quantity of base required to reach the equivalence point is the same in the two titrations.

b. The pH is higher initially in the titration of a weak acid.

c. The pH is higher at the equivalence point in the titration of a weak acid.

d. The pH in excess base is essentially the same for the two cases.

e. In titrating a weak acid, one needs an indicator that changes at a higher pH than for the strong acid titration.

40. a. pH = 1.54 b. pH = 3.30

c. pH = 7.00 d. pH = 10.70

e. pH = 12.74

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UNIT 3 INTRODUCTION TO CHEMICAL THERMODYNAMICS

UNIT OVERVIEW

Total of Periods Allotted: 12 periods

In this unit, a branch of chemistry that deals with the transformation of heat and other forms of energy (chemical thermodynamics) will be studied. In thermodynamics, there are different thermodynamic terms. These are studied in section 3.1. The first law of thermodynamics and some thermodynamic quantities will be studied in section 3.2. Section 3.3 deals with thermochemistry, under which heat of reactions, standard states, Hess's law and bond energies will be studied. Finally, entropy and the second law of thermodynamics will be studied in section 3.4. Similar to other units, different activities, experiments and exercises are also given in this unit to promote the active-learning teaching-learning process.

The methodologies suggested for teaching this section are group discussion, gapped lecture, experiment and question-and-answer.

Unit Outcomes

After completing this unit, students will be able to:

- *understand the terms: system, open system, closed system and isolated system, spontaneous process, state and path function, intensive and extensive property, internal energy, heat and work;*
- *understand the first and second laws of thermodynamics;*
- *compare the energy changes observed when chemical bonds are formed and broken, and relate these changes to endothermic and exothermic reactions;*
- *identify ways in which the terms reactant, product, and heat are combined to form thermodynamic equations representing endothermic and exothermic chemical changes;*
- *understand concepts such as enthalpy change, entropy change and free-energy changes;*
- *determine the spontaneity of a given reaction;*
- *demonstrate scientific enquiry skills, including: communicating, measuring, applying concepts and asking questions.*

Main Contents

- **3.1 Common thermodynamic terms**
- **3.2 The first law of thermodynamics and some thermodynamic quantities**
- **3.3 Thermochemistry**
- **3.4 Entropy and the second law of thermodynamics.**

3.1 COMMON THERMODYNAMIC TERMS

Periods Allotted: 2 periods

Competencies

At the end of this section, students will be able to:

- *define chemical thermodynamics;*
- *define systems;*
- *give examples of systems;*
- *explain open, closed and isolated systems;*
- *classify systems as open, closed and isolated system;*
- *define state function and path function;*
- *distinguish between intensive and extensive properties;*
- *define spontaneous and nonspontaneous process.*

Forward Planning

Thermodynamics, is a new concept for the students. So special preparation is needed to explain the main idea of the section. Common thermodynamic terms need to be elaborated with examples and questions and methodology. Make a special arrangement for the students so that they can do **Activities 3.1**, **3.2** and the start up activity.

Subject Matter Presentation

Implement the group discussion, gapped lecture and brainstorming methodologies to teach this section.

Start with the startup activity. The activity is designed to help students discover the relationships between energy transformations they encounter in their daily lives with the concept of thermodynamics. Let the students discuss Activity 3.1 and present their reports. Harmonize their discussion by describing the importance of thermodynamics in chemistry and other fields such as engineering. Emphasize the fact that thermodynamics deals with energy conversion and the stability of molecules and direction of change. To explain the importance of thermodynamics to society, you can give the following example. Some of the largest industries manufacture products that release, absorb or limit the flow of energy. Common fuels (oil, wood, coal and natural gas) release energy for heating and for powering combustion engines and steam turbines. Fertilizers help crops absorb solar energy and convert it to the chemical energy of food, which our bodies convert into other forms. Numerous plastic, fiberglass, and ceramic materials serve as insulators that limit the flow of energy. Hydroelectric power is used as a source of electricity. Mention that these applications are widely used in Ethiopia.

Following this, present Activity 3.1. The activity teaches students some common thermodynamic terms. So, let students discuss Activity 3.1 in their groups for a few minutes and then have some groups present their opinions to the class. After the presentations, harmonize concepts.

Due emphasis should be given to the fact that, in studying thermodynamics, terms such as **the system** and **surroundings** should be considered. In order to observe and measure a change in energy, we must first define the system, which is the part of the universe that we are going to focus on. The moment that we have defined the system, everything else relevant to the change is defined as the surroundings.

Systems may be open, closed or isolated, depending on the types of exchange occurring between the system and surroundings. Let the students give more examples of the three types of systems.

Define a state and explain the difference between state function and path function. Properties of the system such as intensive properties and extensive properties should be described. The students should know what spontaneous process and nonspontaneous process are and be able to give examples of each process.

Let the students list the properties of a system and classify them as extensive and intensive (**Activity 3.2**). Here, you can apply the brainstorming method. After students have completed their work on Activity 3.2, let the individual groups each suggest one property as intensive or extensive. Record all the properties they suggest on the blackboard. After their attempt, you can harmonize concepts and give them the following list of properties.

Properties	Intensive	Extensive
Color	√	
Length		√
Mass		√
Density	√	
Volume		√
Melting point	√	
Boiling point	√	
Ductility	√	
Malleability	√	

The measured value of an extensive property depends on how much matter is being considered. Mass, which is the quantity of matter in a given sample of a substance, is an extensive property. More matter means more mass. I.e. the value of an extensive quantity depends on the amount of matter. The measured value of an intensive property does not depend on how much matter is being considered. Density, defined as the mass of an object divided by its volume, is an intensive property. So is temperature.

Assessment

To assess each student's work in relation to common thermodynamic terms (section 3.1), record the participation of every student in discussing the start up activity and Activities 3.1 and 3.2.

Ask the students to read and understand the meanings of the following terms and write the definitions in their own words: thermodynamics, system, surroundings, isothermal, adiabatic, state, equation of state, heat capacity, specific heat, state function and state variable. It would be good if they try to explain the terms to their classmates verbally.

From what you observe about their performances, check whether or not the suggested competencies for this section are achieved by most of the students.

Additional Questions

* 1. identify the following as an open, closed and isolated system.
   a. A bottle of Coca Cola kept in a refrigerator.
   b. A chemical reaction taking place in a thermos flask.
   c. A mixture of zinc and hydrochloric acid reacting in a beaker.
   d. A sample of calcium carbonate in a crucible that decomposes by heat.

* 2. What is the difference between state function and path function?

Answers to Additional Questions

1. a. closed system
   b. isolated system
   c. open system
   d. open system
2. State function is a thermodynamic quantity that depends on the initial and final states of the system. Path functions are quantities that do not depend on the initial and final states of the system but depend on the path followed to prepare a particular state of a system.

3.2 FIRST LAW OF THERMODYNAMICS AND SOME THERMODYNAMIC QUANTITIES

Periods Allotted: 3 periods

Competencies

At the end of this section, students will be able to:

- *explain internal energy in relation to the concepts of thermodynamics;*
- *explain heat in relation to the concepts of thermodynamics;*
- *explain work in relation to the concepts of thermodynamics;*
- *state the first law of thermodynamics;*
- *explain the first law of thermodynamics;*
- *calculate the change in the internal energy of a system, based on a given information.*

Forward Planning

Read about and understand internal energy, heat, work and the first law of thermodynamics from the students' textbook and make the necessary preparations to explain them.

Have a look at the examples given in the students' textbook and prepare additional sample problems on internal energy, heat and work, and solve these problems yourself before you introduce them to the class. Read the activities given in this section beforehand. Plan how to group the students to do the activities.

Make a plan of your own that shows the topics and activities you are going to deal with during each period in order to complete the whole section within three periods. Indicate the time allotted for each activity you perform in each period.

Subject Matter Presentation

Internal Energy, Heat and Work

To teach this section, use group discussion, question-and-answer as methodologies.
Start the discussion on *Internal Energy (E)* using Activity 3.3. The activity is designed to help student discover that pumping air into a bicycle tire results in work done on the gas, an increase in energy and also an increase in temperature. Let students discuss the activity, in their groups, for a few minutes. Then have some groups suggest their opinions to the class.
Then, harmonize concepts as follows. The warming effect at the valve stem during the pumping of air into a bicycle can be explained by the first law of thermodynamics. The

action of the pump compresses the air inside the pump and the tire. The process is rapid enough to be treated as approximately adiabatic, so that $q = 0$ and $\Delta \in = w$. Because work is done on the gas in this case, w is positive, and there is an increase in energy. Hence, the temperature of the system also increases.

Each particle in a system has potential and kinetic energy, and the sum of these energies for all the particles in the system is the internal energy, E. When a chemical system changes from reactants to products and the products return to the starting temperature, the internal energy has changed. To determine this change, ΔE, we measure the differences between the system's internal energy after the change (E_{final}) and before the change (E_{inital}).

$$\Delta E = E_{\text{final}} - E_{\text{initial}} = E_{\text{products}} - E_{\text{reactants}}$$

Inform students that kinetic energy is a component of internal energy consisting of various types of molecular motion and the movement of electrons within molecules. Let them also know that potential energy may result from attractive interactions between molecules, between nuclei and electrons within molecules or repulsive interactions between electrons within molecules or between nuclei within molecules.

Then, continue by presenting activity 3.4. The activity is designed to help students realize that the absorbation of different amounts of heat energy by two identical substances of equal mass does not result in the same amount of temperature increase.

Have students do Activity 3.4 then in groups for a few minutes. After they complete the discussion, let one or two groups present their conclusions and opinions to the class.

After the presentation, tell them that the 150W bulb produces higher energy than the 100W bulb. So the temperature of the paper placed on the 150W will be higher than that of the paper on the 100W bulb.

After harmonizing concepts, continue by teaching about heat. Explain the distinctions between different forms of energy. One distinction that is made is between heat and work. Heat is the exchange of thermal energy from a hot body to a cold body. Explain to the students that heat is a mode of transfer of energy when two bodies at different temperature are placed in contact. Heat flows from a body of higher temperature to one with lower temperature.

Then, tell the students about the method of measuring the quantity of heat flow. Define terms like calorimetry, calorimeter, heat capacity and specific heat. Solve some problems to show your students how they can calculate heat absorbed or released during a chemical reaction or physical process. Also, give them some exercises and check their work.

After that, proceed to Activity 3.5. This activity enables students to find out whether or not two samples of the same substance of equal mass can absorb or release the same amount of heat energy in the same interval of time. So let students discuss Activity 3.5 in groups for a few minutes and then have some groups present their opinions to the class. After the presentations, harmonize concepts as follows.

The amount of water in the two beakers is the same and is of equal mass. In addition to this, the change in temperature is also the same. Since the amount of energy released or absorbed is determined by $q = mc\Delta T$, the amount of heat energy absorbed by the water whose temperature was raised by 10°C is exactly equal to the heat released by the water whose temperature was reduced by 10°C. Thus, the water samples in both beakers attain room temperature in the same interval of time.

After harmonizing concepts for Activity 3.5, continue by teaching about work. First tell them that work is a path function. Let them also know that, in this part, we give present only the pressure - volume work of a gas. Introduce them to the formula used in calculating pressure - volume work. Also explain the signs of work done during the expansion and compression of gases.

Finally, proceed to Activity 3.6, which helps students to identify different types of energy transformations at home or in their environment. Let them discuss this activity in groups for a few minutes and then have one or two students from different groups present their opinions to the class. After the presentations, harmonize concepts.

You can mention that electric energy is converted into heat and light energy at home in light bulbs and heaters. The energy obtained by the combustion of fuels can be used for heating, driving vehicles etc. In dry cells and batteries chemical energy is converted to electrical energy.

The First Law of Thermodynamics

Start the lesson by asking students to state the law of conservation of energy. After their responses, tell them that the first law of thermodynamics is a restatement of the law of conservation of energy. Let them know the mathematical expression of the first law:

$$\Delta E = q + w$$

Let them also get information about the sign of ‘q’ when heat flows into and out of the system and also the sign of ‘w’ when work is done on the system by the surroundings and when the system does work on its surroundings.

Solve some problems as examples to show the application of the first law of thermodynamics and let students practice solving problems in relation to this law. Give them exercise as class work or homework.

Assessment

Assess each student's work throughout section 3.2, which presents the first law of thermodynamics and some thermodynamic quantities. To do this, record how every student is doing in your permanent performance list. Your record may be based on how each student is taking part in

- discussing Activities 3.3 – 3.6
- presenting the views, opinions and conclusions of the group to the rest of the class after discussion
- answering questions raised during discussions, the process of harmonizing concepts, mini-lectures and the process of stabilization

Also, give them Exercises 3.1 and 3.2 as class work. Check their work and record their achievements.

From your record, make sure that the suggested competencies for section 3.2 are achieved by most of the students. Praise students working above the minimum requirement level and encourage them to continue working hard. With regard to those working below the minimum requirement level, either arrange extra lesson time or give them additional exercises. You may use those recommended to them from the additional questions given in the guide (section 3.2).

Additional Questions

* 1. If 500 cm^3 of a gas is compressed to 250 cm^3 under a constant external pressure of 2.96 atm, and also absorbs 12.5 kJ, what is the change in internal energy?

* 2. The work done when a gas is compared in a cylinder is 299 J. During this process, 7.0 J of heat is transferred from the gas to the surroundings. Calculate the change in the internal energy of the gas.

Answers to Additional Questions

1. $\Delta E = 12.6$ kJ
2. $\Delta E = 229$ J

Answers to Exercise 3.1

a) q = 79,420 J or 79.420 kJ

b)Molar heat capacity = 75.2 J/mol.k

Answer to Exercise 3.2

a. $\Delta E = 77J$ b. $\Delta E = 8 \times 10^7$ J

3.3 THERMOCHEMISTRY

Periods Allotted: 4 periods

Competencies

At the end of this section, students will be able to:

- *define enthalpy change, ΔH;*
- *derive an expression for the enthalpy change of an ideal gas, based on the first law of thermodynamics;*
- *calculate enthalpy change for ideal gases, from given information;*
- *define standard state;*
- *define standard molar enthalpy of combustion,* ΔH_c^o *;*
- *describe how heats of combustion can be used to estimate the energy available from foods;*
- *define standard molar enthalpy of formation,* ΔH_f^o *form;*
- *define standard molar enthalpy of neutralization,* ΔH_{neut}^o *;*
- *carry out an activity to measure standard molar enthalpy of neutralization;*
- *state Hess's law;*
- *apply Hess's law to solve problems on enthalpy changes of chemical reactions;*
- *explain bond energy;*
- *calculate the bond energies of substances in a given chemical reaction;*
- *calculate the standard enthalpy changes of reaction from given enthalpy changes of reactants and products.*

Forward Planning

Read the information on thermochemistry thoroughly and make the necessary preparations to discuss enthalpy, enthalpy change (ΔH), endothermic and exothermic processes, standard states, standard molar enthalpy of formation and Hess's law.

Design a plan that shows the contents and activities you are going to treat during each period in such a manner that the entire contents of the section will be covered with in four periods. In your plan, indicate the duration of time required for every activity you perform in each period.

Prepare chemicals, apparatuses and other materials required to perform Experiment 3.1.

Teaching Aids

Refer to the students' text for the chemicals and apparatus required to perform Experiment 3.1.

Subject Matter Presentation

Heat of Reactions

You may use the group discussion and question-and-answer methods to teach the contents in this topic.

Start the discussion by asking students to define thermochemistry. After their responses, tell them that thermochemistry is the branch of thermodynamics that deals with the heat involved in chemical and physical change, and especially with the concept of enthalpy. Relate the definition to **Activity 3.7.** The activity is designed to assist students to realize the concept of enthalpy by associating it with changes they encounter in their daily lives.

So, let them discuss Activity 3.7 in groups for a few minutes and then have some groups present their ideas. After their presentations, harmonize their discussion. Let them understand that water does not freeze at room temperature because it needs less temperature than the room temperature. The ice melts at room temperature due to an increase in enthalpy.

Using the brief derivation given in the students textbook, show the students that enthalpy, H, is a combination of the terms **$E + pV$.** Let the students notice that the enthalpy change, ΔH, is equal to the heat absorbed under constant pressure when the only form of work is pressure-volume work. Finally ask them the difference between ΔH and ΔE. After their responses, inform to them that $\Delta H = q_p$ while $\Delta E = q_v$

Standard States

You may use the question-and-answer method to teach the contents in this topic. Start the lesson by asking students what the standard state of a substance is. Following their responses, explain the standard state as the normal chemical form of the compound at 1 atm pressure and a specified temperature, usually 298 K. Next, ask them to define standard enthalpy of combustion. Invite one or two students to give the answer. After their attempts, tell them that standard enthalpy of combustion is the enthalpy change that accompanies the combustion of one mole of a substance in oxygen at standard conditions. Give them examples and then continue with standard enthalpy of formation. Before you deal with the details, give the opportunity to the students to guess the definition. After that, tell them the appropriate definition as follows. The standard enthalpy of formation, ΔH_f^o, is the change in enthalpy in the reaction when one mole of a substance is formed from the elements in their standard state. The convention is that the enthalpies of all elements in their standard states are zero. The enthalpy change in any reaction can be calculated by using a table of standard enthalpies of formation (Table 3.2 in the students' textbook).

$$\Delta H^o_{rxn} = \sum n\Delta H^o_{f\,(products)} - \sum m\Delta H^o_{f\,(reac\tan t)}$$

Illustrate this equation with the example given in the students' textbook. Encourage the students to do the exercises given in this section.

Hess's Law

It is preferable to implement analogy and gapped lecture as method teaching for this topic.

For students to realize Hess's law, you can use analogy. For example, a man travelled directly from Addis Ababa to Dire Dawa in one day. At another time, he travelled from Addis Ababa to Adama the first day, from Adama to Asebe Teferei the second day and from Asebe Teferei to Dire Dawa the third day. This man travelled the same distance, whether he completed his trip in a single day or in three days. The distance he covered in his first trip is equal to the sum of the distances he travelled in three days.

The same analogy can be used in the case of Hess's law. The energy change for a reaction is always the same, whether that reaction proceeds in one step or in a series of steps. The energy change for the overall reaction is equal to the sum of the energy changes of the individual steps.

Discuss Hess's Law and illustrate its application, using examples in the students' textbook. To use Hess's law, we imagine an overall reaction as the sum of a series of reaction steps, whether or not it really occurs that way. Each step is chosen because its ΔH is known. The overall ΔH depends only on the initial and final states. Hess's law says that we add together the known ΔH values for the steps to get the unknown ΔH of the overall reaction. Similarly, if we know the ΔH values for the overall reaction and all but one of the steps, we can find the unknown ΔH of that step. Give them one question from Exercise 3.5 as class work. Check their works and give them correction.

Bond Energies

Use gapped lectures and experiment as methodologies to teach the contents in this topic.

Explain to the students that the enthalpy change can be calculated using bond energy values.

$$\Delta H = \sum(\text{bond energies of bond broken}) - \sum(\text{bond energies of bond formed})$$

The bond energies of some substances are given in Table 3.3 in the students' text book. Use the examples given in the textbook and other additional examples to illustrate the bond energy. Let the students do experiment 3.1 to measure the standard molar enthalpy of neutralization of sodium hydroxide and hydrochloric acid. The heat of neutralization can be calculated using the formula $q = mC\Delta T$.

Assessment

You may assess each student's work throughout section 3.3 (Thermochemistry) by recording in your permanent performance list how the student is doing. You can do this by taking into consideration how the student

- ✓ participates in discussing Activity 3.7
- ✓ participates in presenting opinions of the group after discussion
- ✓ answers questions raised during harmonizing concepts, stabilization or mini-lectures
- ✓ performs activities given between gapped lectures
- ✓ participates in performing Experiment 3.1
- ✓ takes part in presenting the observation of the group to the rest of the class

In addition to that, give them Exercise 3.3 – 3.6 as class work or homework. Check their work and record their achievements. Based on the cumulative record you have, make sure that the competencies for this section are achieved by most of the students. Praise students working above the minimum requirement level. Give the necessary assistance to those working below the minimum requirement level, either by arranging extra lesson time or by giving them additional exercises. You may use the Additional Questions given in this guide.

Additional Questions

* 1. Calculate the standard enthalpy change for the following reaction:

$$10N_2O(g) + C_3H_8(g) \rightarrow 10N_2(g) + 3CO_2(g) + 4H_2O(g)$$

The standard enthalpies of formations for N_2O (g), C_3H_8 (g) CO_2 (g) and H_2O (g) are 82, —104, —393.5 and —242 kJ/mol respectively.

2. Calculate ΔH° for the reaction:

$$2N_2(g) + 5O_2(g) \rightarrow 2N_2O_5(g)$$

based on the following information

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell), \Delta H^o = -571.5 kJ$$

$$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(\ell), \Delta H^o = -76.6 kJ$$

$$\frac{1}{2}N_2(g) + \frac{3}{2}O_2(g) + \frac{1}{2}H_2(g) \rightarrow HNO_3(\ell), \Delta H^o = -174.1 kJ$$

* 3. The bond energies of H – H, H – F, N – H and O – H bonds are 436, 565, 389 and 463 kJ/mole, respectively. Which of these bonds is a) the strongest, b) the weakest bond?

4. If the bond energy of O = O bond is given as +498 kJ/mol or –498 kJ/mol, what does the negative or the positive sign indicate?

Answers to Additional Questions

1. ΔH_{rxn} = –2864.5 kJ
2. Reverse the first reaction, multiply the second equation by two and then reverse, multiply the third equation by four and then ΔH = +28.3
3. a. H – F bond b. N – H bond
4. Bond forming is an exothermic process and bond breaking is an endothermic process. Thus bond breaking energy has a positive value and bond forming energy has a negative value. However, bond breaking energy and bond forming energy for the same bond are equal in magnitude but opposite in sign. Thus +498 kJ is the energy required to break the O = O bond while –498 kJ is the energy released per mole of O = bond formed.

Answer to Exercise 3.3

a. ΔH = –890 kJ

b. 1g C_3H_8 produces the heat = –50.5 kJ/g, while 1g C_6H_6 produces the heat

= –41.9 kJ/g

Answer to Exercise 3.4

26.3°C

Answer to Exercise 3.5

1. ΔH_{rxn} = -297 kJ
2. ΔH = 226.7 kJ

Answer to Exercise 3.6

a. ΔH = -15 kJ b. ΔH = -36 kJ

3.4 ENTROPY AND THE SECOND LAW OF THERMODYNAMICS

Periods Allotted: 3 periods

Competencies

At the end of this section, students will be able to:

- *explain entropy and entropy change;*
- *calculate entropy change from the given standard entropies of substances;*
- *state the second law of thermodynamics;*
- *explain the second law of thermodynamics;*
- *calculate the entropy changes from the given enthalpy changes of the system and absolute temperature;*
- *explain free energy and free-energy change;*
- *calculate the standard free-energy change, ΔG^o, from given standard free energies of reactants and products;*
- *describe the relationship between ΔG^o, ΔH^o and ΔS^o, of a reaction;*
- *determine the spontaneity of a given reaction.*

Forward Planning

Read the students' text book and other reference books to understand entropy, entropy changes, standard entropy changes, the second law of thermodynamics, free energy and free energy changes. Prepare yourself for discussing how ΔH, ΔS and ΔG are used to predict whether the process is spontaneous or not.

Make a plan of your own that shows which contents and activities you are going to deal with during each period in order to cover the entire contents within three periods. Read the teacher's guide to get information about the methodologies you implement and also to gain more ideas about Activities 3.8 – 3.10.

Subject Matter Presentation

Entropy and Spontaneous Process

You may use group discussion analogy and question-and-answer as methodologies to teach this topic.

Start the discussion by reviewing how reaction spontaneity can be determined, using enthalpy change of a reaction. Explain to the students that decrease in enthalpy of a system by itself alone is not enough to be taken as an indicator of spontaneity. Spontaneity is favored if there is an energy decrease (ΔH is negative) corresponding to an exothermic change, and also if there is an increase in the degree of randomness or disorder of a system.

To help students realize what an entropy is, take three boxes containing twenty spherical balls of equal size, one box containing blue, the second containing yellow and the third red balls. First show them that in each box, the balls have an ordered arrangement. Then, combine the three sets of balls in to one box and mix them very well. Ask them whether or not the red balls appear one after the other as they did before mixing with others.

After their responses, tell them that the balls are more disordered after mixing them with others than they were before. The random arrangement of the red, yellow and blue balls increased after mixing. Now, ask the students to define entropy in analogous to the above example. Finally, define entropy (S) as a thermodynamic quantity related to randomness or disorder. Spontaneity is favored by an entropy increase (if ΔS is positive). To help students discover factors that make a process spontaneous, let them discuss Activity 3.8 in groups for a few minutes and then have some groups present their findings to the class. After their presentations, harmonize their discussion based on the following points. In general entropy increases for the following processes:

- ✓ solids melt to liquids
- ✓ solids or liquids vaporize to form gases
- ✓ solids or liquids dissolve in solvents to form solution.
- ✓ a chemical reaction produces an increase in the number of molecules of gases.

Finally, introduce students to the formula used in calculating entropy change, ΔS, for a chemical change and let them know what the value of ΔS indicates when it is greater and less than zero. Solve some problems to show students how they can calculate entropy change for a chemical reaction. Let them also practice solving ΔS value for a reaction. Give them exercises as class work or homework.

The Second Law of Thermodynamics

You may implement group discussion and gapped lectures as methods of teaching for this part.

Start the lesson with Activity 3.9. The activity enables students to understand that the universe should be at equilibrium for living things to continue to survive. Have students discuss Activity 3.9 and present their opinions to the class. After their presentations, harmonize concepts. Tell them that living things would be damaged by the solar energy if the process of the universe is not at equilibrium.

Explain the second law of thermodynamics in such a way that it expresses the notion that there is an inherent direction in which any system moves if it is not at equilibrium. For example, if you drop a brick, it falls to the floor. Water placed in a freezer compartment turns into ice. A tiny nail left outdoors eventually rusts. Each of these processes occurs without outside intervention. Such processes are said to be spontaneous.

Note for the Teacher

Let the students know clearly the difference between reaction spontaneity and reaction rate. Explain to them that the fact that a process is spontaneous does not mean that it will occur at an observable rate. A spontaneous reaction may be very fast, as in the case of an acid-base neutralization or very slow, as in the case of rusting of iron. Thermodynamics can tell us the direction and extent of a reaction, but it can say nothing about its speed. Remind the students that they have learned about reaction rates in Grade 11 (Chemical Kinetics).

Explain to the students that the entropy change in a chemical reaction is given by the sum of the entropies of products minus the sum of the entropies of the reactants.

$$\Delta S^o = \sum nS^o_{(products)} - \sum mS^o_{(reactants)}$$

The very important thermodynamic quantity to be mentioned in this section is Gibbs free energy, G, which is defined as G = H – TS. For a process occurring at constant temperature, the change in free energy is given by the expression $\Delta G = \Delta H - T\Delta S$.

Explain to the students that ΔG is the means for incorporating both ΔH and ΔS into a new quantity that tells us whether a reaction will be spontaneous or not.

From the relation

$$\Delta G = \Delta H - T\Delta S$$

There is a simple relationship between the sign of ΔG for a reaction and the spontaneity of that reaction operated at constant temperature and pressure.

- ✓ If ΔG is negative, the reaction is spontaneous in the forward direction.
- ✓ If ΔG is zero, the reaction is at equilibrium.
- ✓ If ΔG is positive, the reaction in the forward direction is non spontaneous.

ΔG^o (the standard free energy change) can be calculated from tabulated ΔG^o_f values.

$$\Delta G^o_{(reaction)} = \sum n\Delta G^o_{f\,(products)} - \sum m\Delta GF^o_{f\,(reac\tan ts)}$$

Let them do Activity 3.10 as a group assignment and submit it to you for correction. The activity is designed to enable students to decide the spontaneity of a process from the combinations of values of ΔH, ΔS, T and ΔG.

The answers for questions in the activity are as follows:

1. a. T > 350 K or 77°C b. T > 16 K or –257°C

   c. T < 111.4 K or –161.6°C

2. a. ΔS > 0, ΔS = +184.85J/k

   b. ΔG > 0 at low temperature and the reaction is non-spontaneous.

   ΔG < 0 at high temperatures and the reaction is spontaneous.

   c. ΔS < 0, ΔS = –184.85 J/k

   ΔG < 0, at low temperatures and the reaction is spontaneous.

   ΔG > 0, at high temperatures and the reaction is non-spontaneous.

Assessment

To assess each student's work when you deal with the contents on entropy and the second law of thermodynamics (section 3.4), you need to follow-up the performance of every student. You may do this, by watching closely how the student

- ✓ participates in discussing Activities 3.8 – 3.10
- ✓ takes part in presenting the opinions of the group to the rest of the class
- ✓ answers questions raised during discussion, harmonizing concepts, mini-lectures and stabilization
- ✓ attempts activities given between gapped lectures

Give them Exercises 3.7 and 3.8 as homework. Check their work and record their achievements. Based on the record about their performances, make sure that the suggested competencies for this section are achieved by most of the students.

Additional Questions

* 1. What does an entropy of a system measure?

* 2. What is the sign of the entropy change for the following process?

   a. Vaporization of methanol
   b. Sublimation of iodine
   c. Dissolution of NaCl in water
   d. Formation of lead sulfate precipitate from a solution of lead nitrate and copper sulfate.

e. Diffusion of one gas into another

f. $2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g)$

g. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(\ell)$

h. $N_2O_4(g) \rightarrow 2NO_2(g)$

Answers to Additional Questions

1. What does an entropy of a system measure?
2. What is the sign of the entropy change for the following process?
    a. Vaporization of methanol
    b. Sublimation of iodine
    c. Dissolution of NaCl in water
    d. Formation of lead sulfate precipitate from a solution of lead nitrate and copper sulfate.
    e. Diffusion of one gas into another
    f. $2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g)$
    g. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(\ell)$
    h. $N_2O_4(g) \rightarrow 2NO_2(g)$

Answers to Exercise 3.7

1. a. Increase, positive b. Decrease, negative

   c. Decrease, negative d. Increase, positive

2. a. Negative b. Negative c. Positive

Answers to Exercise 3.8

1. The reaction proceeds spontaneously in the forward direction but spontaneously in reverse direction at higher temperature.

2. The reaction is nonspontaneous in the forward direction at all temperatures.

3. The reaction will proceed spontaneously in the forward direction at high temperature but is spontaneous in the reverse direction at lower temperatures.

4. The reaction becomes spontaneous in the forward direction at some very high temperature. ($\Delta G_{r \times n} = +\ 488.6$ kJ)

Answers to Review Questions and Problems

Part I

1. B 2. C 3. A 4. D 5. B
6. A 7. A 8. C 9. A 10. D

Part II

11. Refer to the students' textbook.

12. a. Work is a force applied over a distance.

    b. The amount of work done is the magnitude of the force times the distance over which it is applied: $W = F \times d$

13. $\Delta E = 130$ J

14. 70153.8J

15. a A state function is a property that depends only on the physical state

    (pressure, temperature, etc.) of the system, not on the route used to get the current state.

    b. Internal energy, E, is a state function, work is not a state function.

    c. Temperature is a state function regardless of how hot or cold the sample has been. The temperature depends only on its regent condition.

16. Refer to the students' textbook

17. a. 4.9×10^5 J

18. a. 4.18 J/g.K b. 355 kJ

19. ΔH rxn = -297 kJ

20. ΔH rxn = -1170 kJ

21. Refer to the students' textbook.

22. Solid, liquid, gas.

23. a. Increases b. Increases c. Increases d. Decreases

24. $\Delta H = 268$ kJ/mol

25. a. At all temperatures.

    b. Below 1114.3k or 841.3°C

26. $\Delta G^o = -817$ kJ

27. $\Delta H_f^o = -238.7$ kJ

28. $\Delta H_f^o = 0.30$ kJ

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UNIT 4 ELECTROCHEMISTRY

UNIT OVERVIEW **Total Periods allotted: 21 periods**

We know that electrochemistry studies about the inter conversion of electrical energy and chemical energy and is also highly applicable to our daily life. At tertiary education level, it is given as a separate field of study. Even at this introductory level, this Unit is very broad and has five major subunits. The subunits are organized from simple to complex.

The first subunit discusses oxidation–reduction reaction and explains oxidation, reduction, and balancing redox reactions using change in oxidation number and ion-electron methods.

The second subunit focuses on electrolysis of aqueous solutions. This subunit deals with how electrical energy is used to bring about chemical change. To clarify this important concept, the electrolysis of molten sodium chloride and aqueous solutions of different compounds are briefly discussed and some are given in the form of activities.

The third subunit explains the Faraday's laws. In this subunit, the Faraday's first and second laws are briefly discussed using examples. Students should apply these laws by calculating the amount of substance deposited at the cathode and anode.

The fourth subunit discusses the industrial applications of electrolysis. It focuses on how electrolysis contributes to the development of corrosion resistant materials using electroplating. It also lets students explain how very reactive metals, nonmetals and important compounds are produced by electrolysis methods.

The last subunit deals with Galvanic or Voltaic cells. It shows how Galvanic cells convert chemical energy into electrical energy. It also discusses how standard reduction potentials and cell potentials can be determined practically and theoretically. It explains the different types of batteries that are very useful in our daily life. Finally, it introduces corrosion and gives opportunity to the students to discuss mechanisms of preventing corrosion.

In these days, the methods of teaching are shifted from traditional lecture method, which gives little emphasis to students' participation in teaching-learning process, to active learning methods. The key concept in active learning is that *knowledge* is not transferred; information is transferred but *knowledge is created* when a student thinks about the information. Hence in active learning you should provide your students with opportunities to do some work based on the ideas you have given them or from what they already know.

Generally this unit acquaints students with new concepts that they might not encounter in lower grades.

Therefore, active learning methods such as group discussion, gapped lecture, role play, brainstorming, independent work, and other appropriate methods are highly recommended to help the students grasp the concepts of this unit.

Unit Outcomes

After completing this unit, students will be able to:

- *understand the fundamental concepts related to oxidation-reduction reactions;*
- *know the application of redox reactions in the production of new substances and electrical energy;*
- *demonstrate an understanding of fundamental concepts related to the interconversion of chemical and electrical energy;*
- *understand the difference between metallic conduction and electrolytic conduction;*
- *identify and describe the function of the components of Electrolytic and Galvanic cells;*
- *understand the difference between electrolytic and Galvanic cells;*
- *know how to solve problems based on Faraday's first and second laws;*
- *measure, through experimentation, the mass of the metal deposited by electroplating (example copper to copper (II) sulphate), and apply Faraday's Law to relate the mass of the metal deposited to the amount of charge passed;*
- *predict spontaneity of redox reactions and overall cell potentials by studying a table of half-cell reduction potentials;*
- *determine the emf of an electrochemical cell, experimentally or from given data;*
- *explain the application of electrochemistry in our daily lives and in industry;*

- *describe examples of common Galvanic cells and evaluate their environmental and social impact;*
- *describe scientific enquiry skills: observing, classifying, comparing and contrasting, communicating, asking questions, measuring, relating cause and effect and problem solving*

Main Contents

4.1 Oxidation -Reduction Reactions

4.2 Electrolysis of Aqueous Solutions

4.3 Quantitative Aspects of Electrolysis

4.4 Industrial Application of Electrolysis

4.5 Voltaic (Galvanic) Cells

4.1 OXIDATION–REDUCTION REACTIONS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *define redox reaction;*
- *define oxidation in terms of electron transfer and change in oxidation number;*
- *define reduction in terms of electron transfer and change in oxidation number;*
- *describe the oxidizing and reducing agents;*
- *identify the species that are oxidized and reduced in a given redox reaction and determine the oxidizing and reducing agents;*
- *balance a given redox reaction using change in oxidation number method;*
- *balance a given redox reaction using ion-electron method.*

Forward Planning

Dear colleague, this section requires your prior reading about oxidation-reduction reactions and balancing redox reactions using change in oxidation number and ion-electron methods. The contents of this topic are thoroughly discussed in the student text. So, read the text before you begin this subunit. Make sure that all the necessary materials and chemicals are available for the reaction of iron nail with copper (II) sulphate solution.

Teaching Aids

Small beakers, iron metal or iron nail, and 1M $CuSO_4$.

Subject Matter Presentation

For this subunit, it is possible to use gapped lecture, group discussion, presentation, and demonstration as active learning methods.

You can start this topic by grouping students and letting them discuss the **start-up activity**. The start-up activity helps them to define oxidation-reduction reactions, and to identify redox reactions occurring in their body and the surrounding. In addition, it motivates them to associate redox reactions with their daily life. You can facilitate their discussion by clarifying misconceptions (if any). After their discussion, allow them to share their ideas with other groups.

The answers for the start-up activity questions are:

1. Redox reactions involve loss and gain of electrons simultaneously by atoms, molecules or ions.
2. Some of the redox reactions occurring in our body, and the surrounding are: respiration, photosynthesis, corrosion, decay and combustion.

Since students have learnt redox reactions in Grade 10 Chemistry, they are expected to define redox reactions and give examples of redox reactions.

At this stage, students are familiar with respiration (oxidation of glucose) in biology lessons and they may give it as an example. Ask them to identify the oxidizing and reducing agents in respiration.

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

In respiration, glucose is oxidized (reducing agent) and oxygen is reduced (oxidizing agent).

Alternatively, you can give brief lecture about redox reactions for fifteen minutes (gabbed lecture) and then give them the chance to list down redox reactions from their daily life.

Encourage them to form groups and do **Activity 4.1.** This activity is intended to help students to:

1. define basic terminologies like oxidizing agent, reducing agent and oxidation number,
2. explain the oxidation of zinc metal and reduction of copper (II) ions in a reaction,

3. distinguish redox and non-redox reactions. identify oxidized substance(reducing agent) and reduced substance(oxidizing agent) and;
4. calculate the oxidation state of each atom in a given chemical formula.

Inform them to assign group representatives for class presentation. You can help them by facilitating group presentations and giving feedback to the class. Try to harmonize their views by providing the following answers of **Activity 4.1**

1. a. Oxidizing agent is a species being reduced and causes oxidation.
   b. Reducing agent is a species being oxidized and causes reduction.
   c. The oxidation number or oxidation state of a bonded atom is the charge it would have if all the electrons in each bond were given to the more electronegative atom.
2. When a piece of iron nail is placed into a blue copper sulphate solution, after some minutes the blue solution fades and a reddish substance is deposited on the iron nail.

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Cu(s) + Fe^{2+}(aq)$$

Black *Blue solution* *Red* *light green solution*

3. The reactions in (a) and (c) are redox reactions . In reaction (a), Ca is oxidized and hydrogen is reduced, and in reaction (c), Si is oxidized and Cl_2 is reduced. However, the reaction in (b) and (d) are non-redox reactions because there is no change in oxidation state.
4. a. Mg b. Zn^{2+} c. Zn^{2+} d. Mg
5. a. +3 b. +6 c. +5 d. +6 e. +7

If students have difficulties in understanding redox reactions, you can give this opportunity to gifted students to demonstrate to the students about the reaction between iron metal and copper sulphate solution.

Prepare 1 M $CuSO_4$ solution by dissolving 159.5 g of $CuSO_4$ in water until the volume of the solution is 1 litre. Then put about 2 mL of the solution in a small beaker and add iron metal or iron nail (as shown in Figure 4.1) to it. Let students observe the deposition of copper metal on iron.

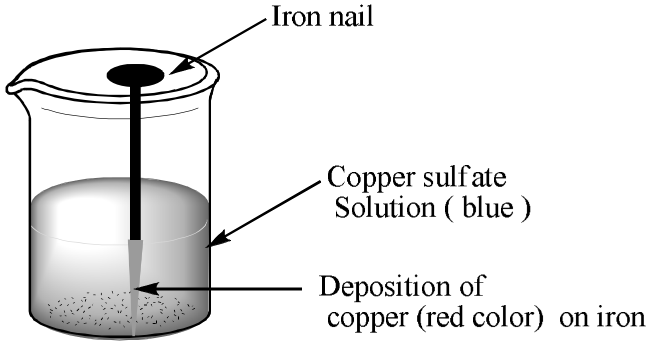


Figure 4.1 Deposition copper metal on iron

Balancing Oxidation-Reduction (Redox) Reactions

For this topic, it is advisable to use gapped lecture, group discussion, and presentation as active learning methods.

This section requires the concepts of the previous section. Therefore, make sure that students do not have problems in identifying redox reactions and in determining oxidation numbers.

The change in oxidation number method requires seven steps. Discuss these steps with them and illustrate the application of these steps using **Examples 4.1 and 4.2**. Then, group the students and give them **Exercise 4.1** to balance the equations. Finally, provide an opportunity to students to show how they balanced the chemical equations on the blackboard.

Similarly, the ion–electron method requires eight steps. Explain these steps to the students and show them how equations are balanced by this method in acidic media using **Example 4.3.** Then, tell students to balance the half reactions in **Exercise 4.2** independently, and finally give a chance to the students to present how they balanced the chemical equations to their classmates on the blackboard.

Give them feedback in each case.

Note for the Teacher

A substance that accepts electrons (*being reduced*) causes another substance to oxidize, and it is called the **oxidizing agent**. A substance that loses electrons (being oxidized) causes another substance to get reduced, and it is called a **reducing agent**.

Here, an agent is a cause for something to occur. If you consider the action of a bleaching agent on clothing in a washing machine, the bleaching agent, which causes a whitening to occur, does not itself become whiter but it causes the cloth to become whiter. Hence, it is possible to conclude that an oxidizing agent is not oxidized and a reducing agent is not reduced.

We find oxidation-reduction reactions in our everyday life. For example, photosynthesis, respiration, corrosion, decay, and combustion are among the many natural processes that involve redox reactions. Redox reaction also occurs in batteries when these are used to provide electrical power, when bleaches are used to remove stains and when antiseptics are used on wounds etc.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓discussing the start-up activity and Activity 4.1
- ✓presenting their views after discussion
- ✓answering the questions in Exercise 4.1 and Exercise 4.2.

By observing their performances from the record, provide them feedback to improve students' learning (formative assessment). In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

*1. When do we say a substance is reduced or oxidized?

2. Is it possible to identify an oxidizing agent and reducing agent? How?

*3. Use the following chemical equation to answer the questions that follow

$$2\,Ag^{+}(aq) \quad + \quad Zn(s) \longrightarrow 2\,Ag(s) + Zn^{2+}(aq)$$

Write the:

a oxidized substance c oxidizing agent

b reduced substance d reducing agent

*4. Balance the reaction in (a) using change in oxidation number method and in (b) using ion-electron method

a. $Cu(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + SO_2(g) + H_2O(l)$

*b. $MnO_4^-(aq) + SO_3^{2-}(aq) \longrightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$ in acidic media

c. $Bi^{3+} + Br_2 \longrightarrow BiO_3^- + Br^-$ *(in basic solution)*

*d $Sn + HNO_3 \longrightarrow Sn(NO_3)_2 + NO_2 + H_2O$

Answers to Additional Questions

1. A substance is reduced if it gains electrons and a substance is oxidized if it loses electrons.
2. Yes it is. By calculating changes in oxidation numbers: If a species decreases its oxidation state in a reaction, it becomes oxidizing agent. Conversely, if a species increases its oxidation state in a reaction, it becomes reducing agent
3. a. Zn b. Ag^+ c. Ag^+ d. Zn
4. a. $Cu(s) + 2H_2SO_4(aq) \longrightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$

   b. $2MnO_4^-(aq) + 5SO_3^{2-}(aq) + 6H^+ \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$

   c. $Br_2 + Bi^{3+} + 6OH^- \longrightarrow 2Br^- + BiO_3^- + 3H_2O$

   d. $Sn + 4HNO_3 \longrightarrow Sn(NO_3)_2 + 2NO_2 + 2H_2O$

Answers to Exercises

Exercise 4.1

a. $Na_2Cr_2O_7 + 6FeCl_2 + 14HCl \longrightarrow 2CrCl_3 + 2NaCl + 6FeCl_3 + 7H_2O$

b. $3Cu + 4HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO_2 + 2H_2O$

c. $3H_2S + 2HNO_3 \longrightarrow 3S + 2NO + 4H_2O$

d. $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + SO_2 + 2H_2O$

e. $P_4 + 20 HNO_3 \longrightarrow 4H_3PO_4 + 20NO_2 + 4H_2O$

Exercise 4.2

a. $2MnO_4^- + 5SO_3^{2-} + 6H^+ \longrightarrow 2Mn^{2+} + 5SO_4^{2-} + 3H_2O$

b. $3Sn^{2+} + 2Bi^{3+} + 18OH^- \longrightarrow 3 [Sn (OH)_6]^{2-} + 3Bi$

4.2 ELECTROLYSIS OF AQUEOUS SOLUTIONS

Periods Allotted:5 periods

Competencies

At the end of this section, students will be able to:

- *recall metallic conductivity, electrolytic conductivity and electrolysis of molten electrolytes;*
- *draw and label diagrams of electrolytic cell;*
- *define preferential discharge;*
- *explain factors that affect preferential discharge;*
- *describe the effects of nature of ions on the electrolysis of aqueous solutions of dilute.* H_2SO_4 *and NaOH;*
- *describe the effect of concentrations of ions on the electrolysis of dilute and concentrated solutions of NaCl;*
- *describe the effects of types of electrode on the electrolysis of* $CuSO_4$ *solution using Ni and Cu electrodes;*
- *write the electrode half reactions and overall reactions for the electrolysis of dilute* H_2SO_4 *, NaOH, dil. NaCl, concentrated NaCl and* $CuSO_4$ *solutions;*
- *distinguish between molten and electrolytes and aqueous electrolytes.*

Forward Planning

Prior reading about the principle of conduction of metals and electrolytes is very useful. Moreover, before you start the subunit, it advisable to familiarize yourself with preferential discharge, electrolysis of molten and aqueous solution of the different compounds listed in the student text.

A conductivity test (Experiment 4.1) is used to let students verify the conduction of metals and electrolytes. Therefore, make sure that all the following materials and solutions for the experiment are ready. In addition, you need to prepare a diagram that shows the components of an electrolytic cell.

Teaching Aids

copper wire, light bulb, power supply (DC) or dry cells, scotch tape, distilled water, file or sand paper, 1M NaCl, 1 M H_2SO_4, 1 M CH_3COOH, 1 M sugar solutions, magnesium ribbon, an iron nail and electrolytic cell diagram.

Subject Matter Presentation

For this subunit, it is possible to use gapped lecture, group discussion, presentation, and experiment as active learning methods.

Students have learned metallic and electrolytic conductivity in Grade 10 Chemistry. Hence it is advisable to start this subunit by letting them recall what they learned in Grade 10 Chemistry. Encourage them to form groups and discuss the questions in Activity 4.2. Then, allow them to present what they discussed to the class by assigning group representatives. You can help them by facilitating group formations and harmonizing group discussions. The activity helps them to identify the charge carriers in metals and electrolytes. Let them come up with general agreements by providing the following answers for Activity 4.2.

1. Conduction in metals is due to the movement of delocalized electrons, but the conduction of electrolytes is caused by the migration of ions to oppositely charged electrodes.
2. To conduct an electric current, a substance should have either delocalized electrons or free mobile ions.
3. Metals can conduct electric current in all physical states. However, electrolytes conduct electric current in molten states or in aqueous solutions.
4. Ionic compounds do not conduct electric current in solid state because cations and anions are not free to move.

Before students start to conduct **Experiment 4.1,** perform the experiment by yourself and make sure that it gives the expected results. The procedure is indicated in the Student Text. If a material is plugged to complete the circuit of Figure 4.2 (a) and the bulb gives light, the material is conductor. Otherwise it is non-conductor or insulator. It is important to polish the surfaces of magnesium ribbon and iron nail because their surfaces are covered with oxide layer. As a result, the oxide layer precludes electrical conductivity.

To distinguish strong electrolytes, weak electrolytes and non-electrolytes, you need to have two dry cells (here higher potential difference is required to see observable changes). This is because the resistance of an electrolyte solution is higher than a solid conductor and four beakers or they may use one beaker by washing it with water. For this experiment, they need 1 M sodium chloride solution, 1 M sugar solution, 1 M acetic acid solution and 1 M sulphuric acid solution. Students should immerse the open ends of the circuit in each solution.

Ask them to classify these solutions as strong electrolytes, weak electrolytes and non-electrolytes based on the brightness of the bulb. The brighter the bulb, the stronger the electrolyte, and the dimmer the bulb, the weaker the electrolyte. The bulb does not glow at all for non-electrolytes.

Anode is the terminal that attached to the positive terminal of the battery and cathode is the one attached to the negative terminal of the battery. Students usually have difficulties identifying cathodes and anodes practically. This is a great opportunity for the students to internalize the concepts. Tell them to note down their observations on the experiment and write a report.

Then, give **Exercise 4.3** to the students to have them classify substances as insulators and conductors. Students should be aware that metallic conduction is a physical process while electrolytic conduction brings chemical change that occurs on the surfaces of electrodes.

Electrolytic cells

For this subtopic, it is advisable to use group discussion and presentation as active learning methods.

Help students to form groups and discuss the questions in **Activity 4.3**, and after their discussion, encourage them to present their answers to the class. The activity helps them to recapitulate some of the concepts they learnt in Experiment 4.1.

To check their progress, you can use a diagram of an electrolytic cell and ask them to label the components of the cell. Harmonize their views using the following answers for **Activity 4.3**:

1. Ions usually migrate towards oppositely charged electrodes. Hence, cations migrate towards cathode and anions to anode.
2. a. Anode is the electrode that is attached to the positive terminal of the battery while cathode is attached to the negative terminal of the battery.

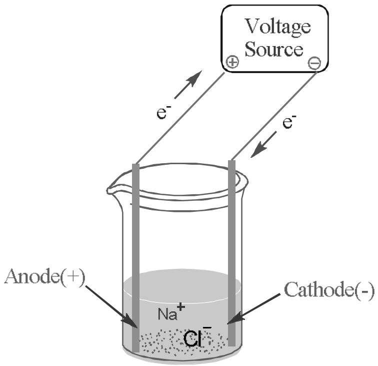


Figure 4.3 (a) Electrolytic cell

b.

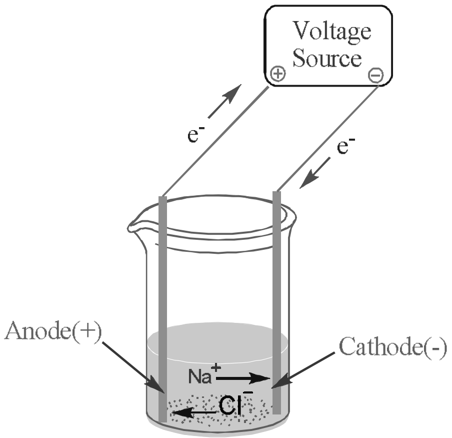


Figure 4.3 (b) Electrolytic cell

c. *Cathode (reduction)*: $2Na^+(l) + 2e^- \longrightarrow 2Na(l)$

Anode (oxidation): $2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$

Overall reaction: $2Na^+(l) + 2Cl^-(l) \longrightarrow 2Na(l) + Cl_2(g)$

d. Sodium metal and chlorine gas.

Preferential Discharge

For this subtopic, it is recommended to use group discussion, presentation and independent work as active learning methods.

The subject of this section may be new for students. Therefore, explain preferential discharge and the three major factors that affect the preferential discharge of ions. Check their progress by giving **Exercise 4.4** to work independently and show their answers to you. Then, give your comments to them.

Students have learned the electrolysis of molten salts in Grade 10 chemistry. Hence, it is advisable to encourage them to do **Activity 4.4** in groups or they can do the activity independently and present their answers to the class. This Activity lets students emphasize preferential discharge of ions. Harmonize their discussion by providing the following answers.

1. Solid sodium chloride does not conduct electricity, because there are no free ions in solid sodium chloride. However, in molten state, there are free ions to conduct electricity.
2. If water is added to salts of active metals, hydrogen gas is reduced at cathode, and oxygen gas is oxidized at the anode.
3. a. at the anode (oxidation): $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$
   b. At the cathode (reduction): $2K^+(l) + 2e^- \rightarrow 2K(l)$
   c. Overall reaction: $2Cl^-(l) + 2K^+(l) \rightarrow Cl_2(g) + 2K(l)$

4. The ease of discharge of H^+ is much greater than Na^+. Similarly, the ease of discharge of OH^- is greater than Cl^- (see Table 4.1 in the Student Text). But if the solution is concentrated, due to high concentration of Cl^- ions, the chloride ions are discharged in preference to OH^- ions.

In the electrolysis of molten ionic compounds, students should be aware that there are only two kinds of ions and we do not need to talk about preferential discharge. However, in the electrolysis of aqueous solutions, the three major factors which affect preferential discharge should be taken into consideration to predict the preferential discharge of ions.

To help students internalize these concepts, the electrolysis of sodium chloride solution, the electrolysis of sulphuric acid solution and the electrolysis of copper sulphate (II) solution are treated.

In the electrolysis of brine solution, tell them to read the content in their text book independently and explain what they read to their classmates. Help them by clarifying concepts that are not clear for them (if any).

Electrolysis of Some Selected Aqueous Solutions

For this subtopic, it is better to use group discussion, presentation and gapped lecture as active learning methods.

In **Activity 4.5**, help them to form groups and then facilitate their group discussions by encouraging each group member to participate actively in his/her group. Then, allow them to present their answers to the class. Praise the presenters and give them feedback. Students are expected to define dilute solution, identify the ionic species in dilute sodium chloride solution, and the reduction half-reaction. Harmonize the students' discussion by providing them the following answers.

1. A dilute solution contains a lesser amount of dissolved solutes.
2. a. Dilute sodium chloride solution contains Na^+, Cl^-, H^+ and OH^- ions.
   b. Na^+ and H^+ ions migrate towards the cathode, and Cl^- and OH^- ions migrate towards the anode.
   c. At the cathode, the H^+ ions are discharged in preference to the Na^+ ions. Since the concentration of the Cl^- ions is very small in dilute sodium chloride solution, the OH^- ions are discharged in preference to the Cl^- ions.

3.

Cathode (reduction): $(2H^+ (aq) + 2e^- \longrightarrow H_2 (g))\times 2$

Anode (oxidation): $4OH^-(aq) \longrightarrow 2H_2O (l) + O_2 (g) + 4e^-$

$$4H^+ (aq) + \cancel{4e^-} + 4OH^-(aq) \longrightarrow 2H_2O (l) + 2H_2 (g) + O_2 (g) + \cancel{4e^-}$$

$$\overset{2}{\cancel{4}}H_2O (l) \longrightarrow \cancel{2}H_2O (l) + 2H_2 (g) + O_2 (g)$$

Overall Reaction: $2H_2O (l) \longrightarrow 2H_2 (g) + O_2 (g)$

In Experiment 4.2, students will electrolyze sodium chloride solution. Hence, help them to construct the electrolytic cell found in the Student Text. Tell them to record their observations and write a laboratory report about the experiment.

Answers to the Questions in Experiment 4.2

1. Bubbling of gases will be observed on the surfaces of the electrodes.
2. The graphite electrode attached to the positive terminal of the battery becomes anode and the other connected to the negative terminal becomes cathode.
3. Hydrogen gas discharges at the cathode and oxygen gas discharges at the anode. However, in concentrated sodium chloride solution, hydrogen gas discharges at the cathode and chlorine gas discharges at the anode.
4. Anode (+) : $4OH^- (aq) + 4e^- \rightarrow O_2(g) + H_2O(l)$

   Cathode (-): $4H^+ (aq) + 4e^- \rightarrow 2H_2(g)$

You can use gapped lecture to teach the electrolysis of dilute sulphuric acid solution and the electrolysis of copper sulphate (II) solution, using inert electrodes. Then, give them Activity 4.6 to discuss in groups or independently and present to the class. This activity is designed to help them to explain inert and reactive electrodes, and discuss the effects of the electrodes in the electrolysis of copper (II) sulphate solution. Harmonize group discussions by providing the following answers.

1. An inert electrode is an electrode that does not affect the product of electrolysis. But an active or reactive electrode is an electrode that affects the product of electrolysis or in the electrolysis reaction.
2. Reactive electrodes.
3. a. Cu^{2+}, SO_4^{2-}, H^+ and

   b. Cu^{2+} ions migrate to cathode, and SO_4^{2-} and H^+ ions migrate to anode.

c. Cu^{2+} ions deposit as copper metal and copper metal oxidizes to produce Cu^{2+} ions.

4.

Anode (oxidation): $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

Cathode (reduction): $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

5. If copper slab attached to anode contains more reactive impurities such as iron or zinc, the concentration of copper sulphate decreases because the more reactive metals are oxidized at the anode, equal moles of copper metal are deposited at the anode. However, if the impurities are less reactive metals like gold and silver, after electrolysis the concentration of copper (II) sulphate solution is almost constant because copper electrode is oxidized at the anode and an equal amount of copper (II) ions are reduced at the cathode.

Conclude this section by reiterating the main points.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing from Activities 4.2 to 4.6,
- ✓ presenting their views after discussion,
- ✓ answering from Exercises 4.3 to 4.5

In addition, you can assess their performance by letting them reflect their laboratory reports for Experiment 4.1 and Experiment 4.2. By observing their performances from the record, provide them feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

1. A teacher adds Na_2SO_4 to water to facilitate its electrolysis in a class demonstration.

   a. What is the purpose of Na_2SO_4?

   b. Why is the water electrolyzed instead of the salt?

2. What property allows copper to be purified in the presence of iron and nickel impurities?

*3. Consider the following general electrolytic cell

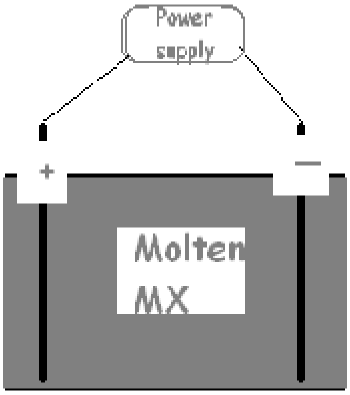


Figure 4.2 Electrolytic cell

a. At which electrode does oxidation occur?
b. At which electrode does element M form?
c. At which electrode are electrons being released by ions?
d. At which electrode are electrons entering the cell?

*4. In the electrolysis of molten NaBr,

a. What product forms at the anode? b. What product forms at the cathode?

In addition, for low achievers identify their learning difficulties and help them until they achieve the minimum required level.

Answers to Additional Questions

1. a. to increase the conductivity of water.
   b. if you look at the preferential discharge of ions: hydrogen ion is discharged in preference to sodium ion and hydroxide ion is discharged in preference to sulphate ion
2. Its ease of discharge or its reduction potential is higher than iron and nickel.
3. a. at positive electrode or anode b. at negative electrode or cathode
   c. at positive electrode d. at negative electrode
4. a. bromine b. sodium

Answers to Exercises

Exercise 4.3

1. a. Iron is a metal and conducts electricity.
   b. Sulphur is a non-metal and is an insulator

c. Solid NaCl does not conduct electricity because there are no free ions.

d. Molten sodium chloride can conduct electricity due to the presence of free ions

Exercise 4.4

1. a. Preferential discharge is the ease in which ions are discharged in preference of other ions during electrolysis.
   b. Inert electrode is an electrode that does not affect the product of electrolysis.
   c. Active or reactive electrode is an electrode that affects the product of electrolysis.
2. Anode (oxidation): $2Cl^{-}(aq) \longrightarrow Cl_2(g) + 2e^{-}$

   Cathode (reduction): $\underline{2H^{+}(aq) + 2e^{-} \longrightarrow H_2(g)}$

   Overall reaction: $2H^{+}(aq) + 2Cl^{-}(aq) \longrightarrow H_2(gl) + Cl_2(g)$

Exercise 4.5

1. a Anode (oxidation): $4OH^{-}(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^{-}$

   Cathode (reaction): $4H^{+}(aq) + 4e^{-} \longrightarrow 2H_2(g)$

   b. Overall reaction: $2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$

2. The concentration of H_2SO_4 increases because water is electrolyzed and its amount decreases.
3. Since copper ions are consumed at the cathode, and hydroxide ions at the anode, the solution becomes acidic due to the formation of H_2SO_4

4.3 QUANTITATIVE ASPECTS OF ELECTROLYSIS

Periods Allotted: 3 periods

Competencies

At the end of this section, students will be able to:

- *state the Faraday's first law of electrolysis;*
- *write the mathematical expressions for the Faraday's first law of electrolysis ;*
- *do calculations related to Faraday's first law of electrolysis;*
- *state the Faraday's second law of electrolysis;*
- *write the mathematical expressions for the Faraday's second law of electrolysis ;*
- *do calculations related to Faraday's second law of electrolysis.*

Forward Planning

Prior reading about the statements of Faraday's Laws of electrolysis and their associated mathematical derivates helps you to identify learning difficulties while conducting the lessons. Before you start this subtopic, look into the contents of the Faraday's laws and the examples from the Student Text.

Subject Matter Presentation

For this subunit, it is possible to use gapped lecture, question and answer, and problem solving as active learning methods.

You can begin this subunit by stating the Faraday's first law and discussing the relationship between charge, current and time (gapped lecture). Show them how the charge of one mole of electrons is calculated and this is the basis for calculating the amount of substance produced during electrolysis.

The steps that help to calculate the amount of substance produced are the following.

Step 1. Calculate the charge passed through the solution ($Q = It$)

Step 2. Change the charge to number of moles of electrons.

Step 3. Convert the mole of electrons derived in step 2 to mole of the element produced at the cathode or anode.

Step 4. Convert number of moles obtained in step 3 to mass in grams.

To understand these four steps, encourage students to carefully look into the examples and identify the steps. Alternatively, they can use the formula: $m = \frac{MIt}{nF}$ where m is mass of the substance deposited or librated, I is current, t is time, F is Faradays constant, n is the number of moles of electrons lost or gained, and M is the molar mass of the substance. However, it is better to use the first method because it helps students to understand the concepts rather than merely memorizing the formula. Let them solve the problems in **Exercise 4.6** and ask willing students to show the solutions of the problems on the blackboard. Praise them and give your comments on their answers.

Similarly, state Faraday's second law, show the derivation of the mathematical equation and let the students observe the **Example 4.6.** Finally, give **Exercise 4.7** as a class work and encourage them to show their solution to the other students. Then, comment on their answers and conclude the session.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student as usual. You can make records based on students' participation in answering the questions in **Exercise 4.6** and **Exercise 4.7.** By observing their performances from the record, provide them feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit as usual.

Additional Questions

1. How long will it take to deposit the cadmium from 300 mL of a 0.4 M $CdSO_4$ solution, using a current of 2 A?
2. What will be the concentration of $Cd(NO_3)_2$ in a solution after 3 A has passed for 5 hr through 900 mL of a solution that was originally 0.3 M (platinum electrodes are used)?

*3. What current is required to plate out 0.020 mol of gold from $AuCl_3$ solution in 3 hr?

*4. How many grams of Na and Cl_2 would be produced if a current of 25 A is passed for 8 hr into molten NaCl?

Answers to Additional Questions

1. 193 min
2. 0.0411 M.
3. 0.536 A.
4. 172 g Na and 265 g Cl_2.

Answers to Exercises

Exercise 4.6

1. Volume of hydrogen is 9.05 L and that of oxygen is 4.52 L. Note that the volume of oxygen is half of hydrogen.
2. Mass of copper deposited at the cathode is 1.78 g and the volume of oxygen gas librated at the anode is 0. 345 L.
3. Mass of Al is 0.672 g

Exercise 4.7

1. The mass of silver is 216 g.
2. Mass of Ag = 21.6 g, mass of Zn = 6.5 g and mass of Fe = 3.73 g

4.4 INDUSTRIAL APPLICATION OF ELECTROLYSIS

Periods Allotted: 2 periods

Competencies

At the end of this section, students will be able to:

- *mention the industrial application of electrochemistry;*
- *explain electroplating and electrorefining;*
- *demonstrate copper refining;*
- *explain how electrolysis is used in the production of some metals, non-metals and compounds.*

Forward Planning

Since this subunit is treated in the form of Activities, prior reading about the industrial applications of electrolysis is very important. Read the note in the Teachers Guide and reference books available in the school library.

Subject Matter Presentation

For this subunit, it is possible to use group discussion, presentation and independent work as active learning methods.

Students learned the application of electrolysis in Grade 10 Chemistry. Let them form groups and do **Activity 4.7.** Help them by facilitating group formations and presentations. Help them to come up with consensus by providing the following answers.

1. Electroplating is the electrolytic deposition of a less active metal on the surface of more active metal to prevent the corrosion of the active metal. Electrorefining is a method of purifying an impure metal using electrolysis.

2. The main purpose of electroplating is to protect metals from corrosion. Sometimes, it is also used to give them a more pleasant or richer appearance. The main purpose of electrorefining is to get rid of impurities from a metal.
3. The anode is often made of the plating metal, while the cathode is usually the metal to be plated.
4. The electrolyte contains the plating metal in the form of dissolved ions

5. Chrome plating is the electrolytic deposition of chromium with other metals. This can usually be done by preparing chromium (VI) solution and by attaching the metal onto which chromium is to be deposited at the cathode.
6. In the electrorefining of copper a thin sheet of pure copper serves as cathode, large slab of impure copper as anode, and aqueous solution of copper sulphate as electrolyte. During electrolysis the impure copper attached to the anode oxidizes and is deposited on cathode. As time passes, the mass of cathode increases and that of anode decreases.

For Activity 4.8 encourage them to work independently and share their answers with their classmates. Harmonize their discussions by providing the following answers.

1. Alkali metals, alkaline earth metals and aluminium are extracted by electrolysis.
2. Since active metals are highly reactive, they cannot be reduced from their compounds using the common reducing agents like carbon.
3. This is due to the fact that molten metal chlorides have lower melting points than other readily available salts.

After they come up with consensus, let them form groups and do **Activity 4.9**. Help them by facilitating group formations and presentations. Tell them to assign group representatives and present to the class. Help them to come up with the following answers.

1. Non-metals such as fluorine, chlorine, hydrogen and oxygen are prepared by electrolysis. Electrolysis is also used in the preparation of compounds such as sodium hydroxide.
2. Anode (oxidation): $2F^-(aq) \longrightarrow F_2(g) + 2e^-$

   Cathode (reduction): $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

   Overall reaction: $2H^+(aq) + 2F^-(aq) \longrightarrow H_2(gl) + F_2(g)$
3. a.

   Cathode (reduction): $2H^+(aq) + 2\cancel{e} \longrightarrow H_2(g)$

   Anode (oxidation) : $2Cl^-(aq) \longrightarrow Cl_2(g) + 2\cancel{e^-}$

   Overall reaction: $2H^+(aq) + 2Cl^-(aq) \longrightarrow H_2(g) + Cl_2(g)$

   The solution becomes basic due to the formation of NaOH.

b.

Cathode (Reduction): $(2H^+(aq) + 2e^- \longrightarrow H_2(g))\times 2$

Anode (Oxidation): $4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$

$$4H^+(aq) + \cancel{4e^-} + 4OH^-(aq) \longrightarrow 2H_2O(l) + 2H_2(g) + O_2(g) + \cancel{4e^-}$$

$$\overset{2}{\cancel{4}}H_2O(l) \longrightarrow \cancel{2}H_2O(l) + 2H_2(g) + O_2(g)$$

Overall reaction: $2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$

4. Pure water is a poor conductor of electricity, so a small amount of ionic solute is added to increase its electrical conductivity.

Conclude this section by recapitulating the main points.

Note for the Teacher

Electroplating is a process in which a metal is electrolytically deposited or "plated out" on the surface of another metal. It is often done to protect the base metal from corrosion or simply to give it a more pleasant or richer appearance. The electrolyte contains the plating metal in the form of dissolved ions. The anode is often made of the plating metal, while the cathode is usually the metal being plated.

One of the most familiar examples of electroplating is "chrome plating" in which a thin layer of chromium is deposited on another metal. The electrolyte is prepared by dissolving CrO_3 in dilute sulphuric acid. The electrolysis reduces chromium (VI) first to chromium (III) and then to chromium metal, the overall reduction being:

$$CrO_3(aq) + 6H^+(aq) + 6e^- \longrightarrow Cr(s) + 3H_2O(l)$$

The chromium deposits on the cathode as a hard protective film.

Electorefining (purification of metals) is another important application of electrolysis. For example, impure copper from the chemical reduction of copper ore is cast into large slabs that serve as the anodes for electrolytic cells. Aqueous copper sulphate is the electrolyte, and thin sheets of ultrapure copper function as the cathode.

The main reaction at the anode is:

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

Other metals, such as iron and zinc, are also oxidized from the impure anode:

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

$$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$$

The Cu^{2+} ions from the solution are preferentially deposited onto the cathode.

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$, producing copper that is 99.95% pure.

The compounds of very active metals are difficult to decompose, so electrolytic reduction is the only method for obtaining such metals. Electrolysis can also be used to prepare less active metals when very pure samples are needed. Molten metal chlorides are usually used as electrolytes, because they generally have lower melting points than other readily available salts. Sodium, lithium, magnesium, calcium, barium and aluminium are among the metals produced by electrolysis.

Electrolysis is also used in the manufacture of non-metals. For example, fluorine, the most electronegative element, is prepared by the electrolysis of a molten mixture of potassium fluoride and hydrogen fluoride. Chlorine is manufactured by the electrolysis of molten sodium chloride or concentrated sodium chloride (*brine*) solution.

Hydrogen and oxygen are also prepared by electrolysis of water. Pure water is a poor conductor of electricity, so a small amount of ionic solute is added to increase its conductivity. The added ions must be more resistant to oxidation and reduction than water.

Electrolysis is also used in the preparation of compounds such as sodium hydroxide.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing the activities from Activity 4.7 to Activity 4.9,
- ✓ presenting their views after discussion,

By observing their performances from the record, provide them feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

1. When a piece of metal A is placed in a solution containing ions of B, metal B plates out on the piece of A.
    a. Which metal is being oxidized?
    b. Which metal is being displaced**?** If bubbles of hydrogen form when B is placed in acid, will they form if A is placed? Explain.
2. How long will it take to plate out each of the following with a current of 100 A?
    a. 1 kg Al from aqueous Al^{3+}.

b. 1 g Ni from aqueous Ni^{2+}.

c. 5 mol Ag from aqueous Ag^{+}.

*3. Which of the following elements are produced primarily by electrolysis?

a. Potassium b. Lead c. Aluminium

d. Zinc e. Calcium

*4. Write the anode and cathode reactions in the electrolysis of copper sulphate using copper electrodes.

Answers to Additional Questions

1. a. A b. B

c. hydrogen bubbles will form when A is placed in acid. Metal A is a better reducing agent than B, so if metal B reduces H^+ in acid, metal will also reduce H^+ ions.

2. a. 30 hrs b. 33 s c. 1.3 hr

3. a, c and e.

4. Anode **:** $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Cathode : $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

4.5 VOLTAIC (GALVANIC) CELLS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *define Voltaic cell and salt bridge;*
- *draw and label Zn – Cu Voltaic cell;*
- *define electrode potential and cell potential;*
- *construct Zn – Cu Voltaic cell;*
- *measure the cell potential of Zn – Cu cell using voltmeter;*
- *explain how standard electrode potential is measured;*
- *describe the reactivity of metal from its position in the activity series or electromotive series;*
- *calculate cell potential;*
- *decide whether a given redox reaction is spontaneous or not;*
- *explain the effect of concentration on cell potential;*
- *mention the different types of Voltaic cells;*
- *give examples of each type of Galvanic cell;*
- *describe how hydrogen-oxygen fuel cell operates;*
- *distinguish between primary and secondary cells;*

- *identify the cathode, anode and the electrolyte of a given Galvanic cell;*
- *compare and contrast electrolytic and Galvanic cell;*
- *explain metallic corrosion in terms of redox reaction;*
- *explain the negative effects of corrosion;*
- *explain the different methods of prevention of corrosion.*

Forward Planning

Dear colleague, this subunit encompasses concepts associated with Galvanic (Voltaic) cells. Before you start to teach the topics, it is better to read all the contents of this subunit from the Student Text book, Teachers Guide, and Reference books available in the school library. Students will determine the Cell potential of Danielle cell (Experiment 4.2). Hence, make sure that all the materials and solutions for the experiment are ready.

Teaching Aids

copper wire, light bulb, power supply (DC) or dry cells, scotch tape, distilled water, a file or sand paper, 1 M NaCl, 1 M H_2SO_4, 1 M CH_3COOH, 1 M sugar solutions, magnesium ribbon and iron nail.

Subject Matter Presentation

For this topic, it is possible to use group discussion and presentation as active learning methods.

Begin this subunit by helping students to form groups and do **Activity 4.10.** Help them by clarifying concepts that are not clear for them while discussing with their groups. In addition, encourage them to share their ideas with other groups. Guide them to come up with the following answers.

1. An electrochemical cell contains two electrodes immersed in their corresponding electrolyte solutions which are connected by a salt bridge. In order to do useful work, the two components of the cell should be separated and connected by a salt bridge.
2. a. Oxidation half-reaction: $Mg(s) \rightarrow Mg^{2+}(aq)$

   b. Reduction half-reaction: $2H^+(aq) \rightarrow H_2(aq)$
3. The reaction in (a) occurs at the anode and the reaction in (b) occurs at the cathode.

Then, tell them to read and discuss the contents of sections 4.5.1 and 4.5.2 with their group members. Facilitate group discussions by clarifying vague concepts for them.

Some students may not be active in group discussions due to various reasons. So, identify these students and encourage them to participate actively in their groups.

Summarize the main points and conclude this session by giving them **Exercise 4.8** as a home work.

Standard Reduction Potentials

For this topic, it is better to use gapped lecture, and experiment as active learning methods.

In this section, use a gapped lecture and discuss how standard reduction potentials are calculated using standard hydrogen electrode (SHE) and explain to them the physical significance or meaning of reduction potentials. If the reduction potential is a large positive value, it is a strong oxidizing agent or the ion is easily reduced. However, when the reduction potential is negative, it has less tendency to be reduced. Instead it has a high tendency to be oxidized and is a reducing agent.

Help them to distinguish the difference between standard reduction potential and cell potential. To clarify these concepts, let students read the examples in their text book and discuss the differences. After finishing their discussion, encourage them to forward their views to the class.

Then, to check their progress, tell them to answer the questions in **Exercise 4.8**. In Experiment 4.3, students will determine the cell potential of Danielle cell. Help them to construct the Galvanic cell found in the Student Text. Tell them to record their observation and write a report about the experiment.

Answers to the Questions in Experiment 4.3

1. The pointer of the ammeter deflects from zinc to copper half cells. This indicates electron flows from zinc to copper half cells. Therefore, copper serves as cathode and zinc serves as anode.
2. (-) Anode (oxidation): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

   (+) cathode (reduction): $Cu^{2+}(aq) \rightarrow Cu(s)$
3. As time passes the intensity of copper sulphate solution decreases or it becomes light blue because the copper ions are consumed in the half-cell.
4. Cell potentials are affected by temperature and concentration.

If the solutions are 1 M, the cell potential becomes exactly 1.10 V.

Electromotive Force

For this topic, it is advisable to use group discussion and presentation as active learning methods.

Students might learn emf and potential energy in Physics lessons. So, encourage them to associate the flow of water from uphill to downhill with emf by doing **Activity 4.11**. So, ask them to explain “what makes electrons to flow in a wire?” Guide them to come up with the following answer.

As it is observed in our daily life, water flows from uphill (high potential energy) to downhill (low potential energy). Similarly, electrons flow in a wire from negative (high electrical potential energy) to positive (low electrical potential energy).

In the spontaneity of reactions section, ask students to explain spontaneous and non-spontaneous processes using the topics they learnt from thermodynamics. Ask them the values of changes in Gibbs free energy for spontaneous and non-spontaneous processes.

Show them the relationship between E^0_{Cell}, ΔG^0 and K by looking at the derivations in the Student Text book and explain to them Example 4.9 and 4.10 that show the calculations of equilibrium constant and change in standard free energy. Finally, let students work **Exercise 4.10** and give feedback for their answers.

In the dependence of emf on concentration section, derive the Nernst equation or help them to derive it, and show how the cell potential is affected by the concentrations of the electrolytes using Example 4.11. Then, allow them to work **Exercise 4.11** and give them feedback.

Batteries

For this topic, it is recommended to use group discussion, presentation and project work as active learning methods.

You can begin this section by letting students do Activity 4.12. The main advantages of batteries are their convenience and portability. It is common to say that a battery is “dead”. A battery is dead when reactants are consumed.

In addition, encourage them to bring “dead” batteries, and classify them as primary, secondary, dry cell etc. Then, ask students to give examples of primary and secondary batteries on the basis of the definitions. The answers of question number 3 depend on batteries collected by students. Tell them ahead of time to go to a nearby automobile garage, and do Project 4.1 and write a report after their visit. This project work shows them the application of electrochemistry in their real life situation.

Students have learned about dry cell batteries in grade 10 chemistry. Hence, help them to form groups and do Activity 4.13. In this activity, students are expected to write out the electrolyte, anode and cathode reactions in a dry cell battery. Harmonize their discussions and guide them to come up with the following answers:

1. The electrolyte consists of ammonium chloride and zinc chloride in water. A carbon rod serves as the cathode and zinc as anode.

2. The cell reactions are

a. $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

b. $2NH_4^+(aq) + 2MnO_2(s) + 2e^- \longrightarrow Mn_2O_3(s) + 2NH_3(aq) + H_2O(l)$

c. $Zn(s) + 2NH_4^+(aq) + 2MnO_2(s) \longrightarrow Zn^{2+}(aq) + 2NH_3(aq) + H_2O(l) + Mn_2O_3(s)$

You can use gapped lecture to teach them alkaline cell, nickel–cadmium cell, and fuel cells. Use the question-and-answer method to engage students in the teaching-learning process.

Then help them to form groups and perform **Activity 4.14.** This activity is designed to help them to associate reversibility of reactions with charging and discharging of batteries. Facilitate their discussion by clarifying vague concepts for them. Help them to come up with the following answers.

Cathode: $PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$

Anode: $PbSO_4(s) + 2H_2O(l) \longrightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

Overall reaction: $2PbSO_4(s) + 2H_2O(l) \longrightarrow Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$

The reactions are the reverse reactions of the discharging of a car battery.

After they come up with consensus for Activity 4.14, let them form groups and do Activity 4.15. Encourage members of each group to participate actively. After discussion, tell them to present their ideas to the class by assigning group representatives.

Harmonize their discussions and guide them to come up with the following answers:

This question is answered using the concept of colligative properties. The concentration of sulphuric acid in a charged battery is usually higher than the discharged battery. The higher the concentration of a solution, the lower its freezing point.

Corrosion of Metals and Corrosion Prevention

For this subunit, it is advisable to use group discussion and presentation as active learning methods.

First, help them to read independently about the corrosion of metals using their Text Book. Then, encourage them to share what they discussed with their classmates. This helps them to associate corrosion with reduction potentials.

Then help them to form groups and discuss the questions in **Activity 4.16.** After discussion, help them to share their ideas with other groups. Help them to reach general agreement by providing the following answers.

1. Painting, galvanizing with zinc, electroplating with less active metals like chromium.
2. All methods prevent the direct contact of iron with oxygen and water.

After they come up with consensus for Activity 4.16, let them do **Activity 4.17** in their groups. Encourage them to present their ideas to the class. Harmonize their discussions and guide them to come up with the following answers:

1. As it is discussed in the note, protection of a metal by making it the cathode in an electrochemical cell is known as cathodic protection. A metal that is oxidized while protecting a metal at the cathode is called the *sacrificial anode.*
2. Only Al, Mn, Mg and Cd can be used as sacrificial anode, because their reduction potentials are less than iron. However, Pb can not be recommended, because its reduction potential is larger than iron and facilitates the corrosion of iron.
3. Magnesium protects iron by being oxidized itself or by serving as sacrificial anode.
4. Anode: $Mg(s) \rightarrow Mg^{2+(}aq) + 2e^-$
5. In cathodic protection, a more active metal than iron is attached to iron pipes. The metal serves as anode and iron becomes cathode. However, less active metals than iron are not recommended for cathodic protection, because in this case iron serves as anode and facilitates the oxidation (rusting) of iron. Students should not advise the replacement of the damaged iron pipe by copper; the reduction potential of copper is +0.15 V which is greater than the reduction of iron (-0.44 V).
6. The reduction potential of zinc is less than iron, and zinc serves as anode and iron serves as cathode in the redox process. Zinc metal scarifies itself by being oxidized and hinders the oxidation (rusting) of iron.

Finally, ask them to submit the answers of the review questions of this unit. You can use the review questions as part of your summative assessment.

Note for the Teacher

The Dry Cell Battery

The most common dry cell, that is, a cell without a fluid component, is the LeClanchè cell used in flashlights and portable radios. It is one of the examples of a primary cell with an irreversible reaction. The cell is not rechargeable.

The anode of the cell consists of a zinc can or container that is in contact with manganese dioxide (MnO_2) and an electrolyte.

The electrolyte consists of ammonium chloride and zinc chloride in water, to which starch is added to thicken the solution to a paste-like consistency so that it is less likely to leak.

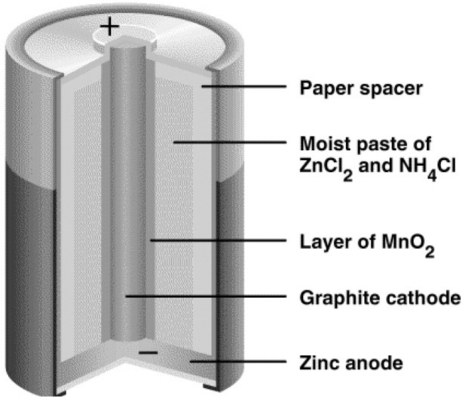


Figure 4.4 The interior section of a dry cell.

A carbon rod serves as the cathode, which is immersed in the electrolyte in the centre of the cell.

The cell reactions are

Anode: $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Cathode: $2NH_4^+(aq) + 2MnO_2(s) + 2e^- \longrightarrow Mn_2O_3(s) + 2NH_3(aq) + H_2O(l)$

Overall reaction: $Zn(s) + 2NH_4^+(aq) + 2MnO_2(s) \longrightarrow Zn^{2+}(aq) + 2NH_3(aq) + H_2O(l) + Mn_2O_3(s)$

The Lead Storage Battery

During the recharging of a lead storage battery, the following reactions occur:

Cathode: $PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$

Anode: $PbSO_4(s) + 2H_2O(l) \longrightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

Overall reaction: $2PbSO_4(s) + 2H_2O(l) \longrightarrow Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$

Note that the overall reaction is exactly the opposite of the normal cell reaction when it is delivering electrical work.

Prevention of Corrosion

The simplest way to prevent corrosion is to ensure that the surface of the metal is not exposed to air and water. Iron is often covered with a coat of paint or another metal, such as tin, zinc or chromium, to protect its surface against corrosion. Tin protects iron only as long as the protective layer is not scratched or damaged. Once it is broken, and the iron is exposed to air and water, tin actually promotes the corrosion of iron. *Why?* If we look at the standard reduction potentials, Fe^{2+} is more difficult to reduce than is Sn^{2+}.

$$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s) \quad E^o_{red} = -0.14\ V$$

$$Fe2+(aq) + 2e^- \longrightarrow Fe(s) \quad E^o_{red} = -0.44\ V$$

As a result, Fe(s) is easier to oxidize than Sn(s). Thus the iron serves as the anode and tin as cathode in the electrochemical corrosion, as shown in **Figure 4.5**

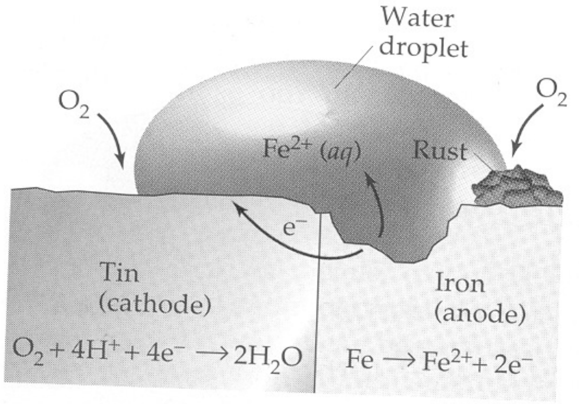


Figure 4.5 Corrosion of iron in contact with tin.

Does the broken layer of zinc coat facilitate the corrosion of iron?

Galvanized iron is produced by coating iron with a thin layer of zinc. The zinc protects the iron against corrosion even after the zinc coat is broken. In this case, the iron serves as the cathode in the electrochemical corrosion, because zinc is oxidized more readily than iron. Therefore, the zinc serves as anode and is corroded instead of iron, as shown in **Figure 4.6**.

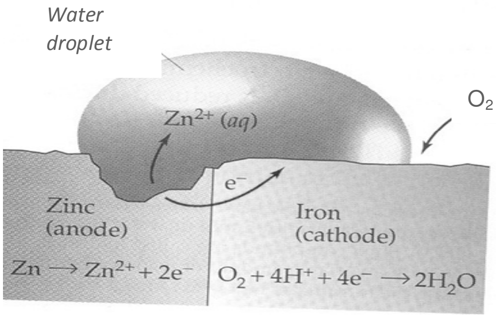


Figure 4.6 Cathodic protection of iron in contact with zinc.

Protection of a metal by making it the cathode in an electrochemical cell is known as cathodic protection. The metal that is oxidized while protecting the cathode is called the *sacrificial anode*. Underground pipelines are often protected against corrosion by making the pipeline as the cathode of a Galvanic cell. Pieces of active metals, such as magnesium, are buried along the pipeline and connected to it by a wire, as shown in **Figure 4.7**.

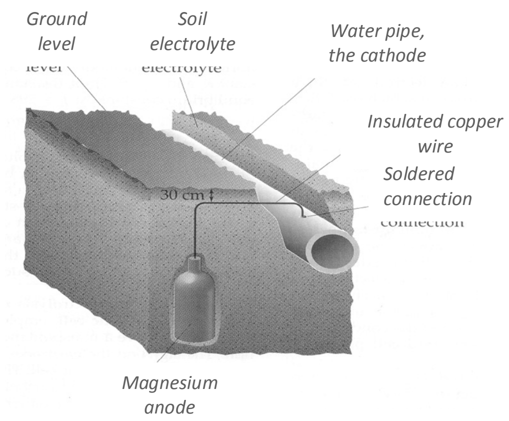


Figure 4.7 Cathodic protection of an iron pipe, using magnesium as the anode.

Assessment

Assess each student's work continuously throughout the subunit as usual. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing during Activity 4.10 to Activity 4.17
- ✓ presenting their views after discussion
- ✓ answering oral questions raised during group discussion.

In addition, give them the exercises from Exercise 4.8 to Exercise 4.13 as class work or home work and record their achievements. By observing their performances from the record, provide them feedback to improve students' learning.

You can also assess their performance by letting them reflect their reports for Experiment 4.3 and project 4.1. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit. You can also use the review questions of this unit as part of summative assessment.

Additional Questions

1. Sketch the galvanic cells based on the following overall reactions. Show the direction of electron flow, the direction of ion migration through the salt bridge, and identify the cathode and anode. Assume that all concentrations are 1 M.

   $Mg + 2Ag^{+} \rightarrow Mg^{2+} + 2Ag$

2. Write the cell notation for question (1).

3. Explain why oxidation and reduction half-reactions have to be physically separated for an oxidation reaction to do work.

*4. In Figure 4.5 magnesium is oxidized and silver ions are reduced

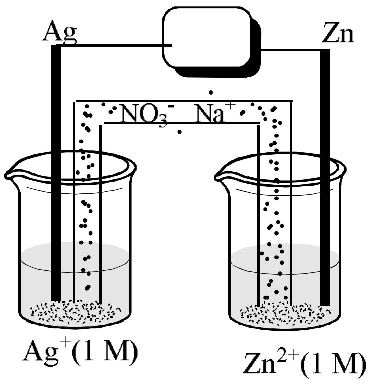


Figure 4.5 Mg-Ag Galvanic cell

a. Label each part of the cell.

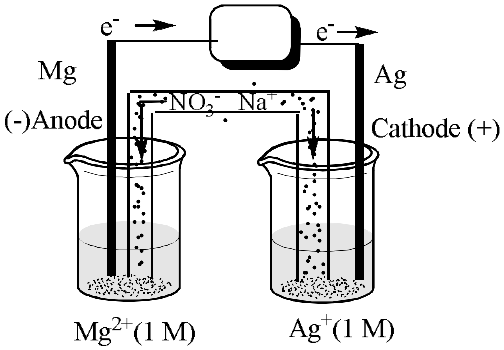
b. Write the cell notation.

c. Write chemical equations for the half-reactions occurring at the anode and cathode, and write an equation for the net reaction in the cell.

d. Show the direction of electron flow and the direction of ion migration through the salt bridge.

Answers to Additional Questions

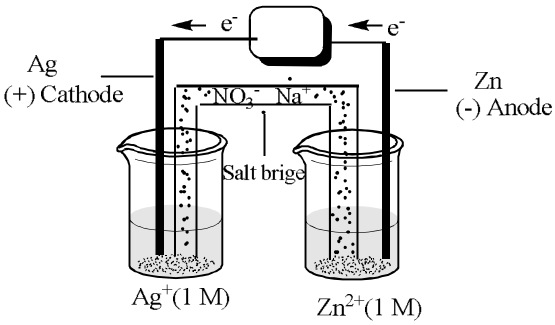
1.



2. Mg (s) | Mg^{2+}(1 M) || Ag^+ (1 M) | Ag(s)

3. If the reactions could occur in one compartment, then the electrons would transfer directly from the reducing agent to the oxidizing agent without doing useful electrical work.

4. a.



b. $Zn\ (s)\ |\ Zn^{2+}(1\ M)\ ||\ Ag^{+}\ (1\ M)\ |\ Ag(s$

c. Anode : $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$
Cathode: $2Ag^{+}(aq) + 2e^{-} \rightarrow 2Ag(s)$
Overall reaction: $Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$

d. Nitrate ions move from salt bridge and enter into zinc compartment, and sodium ions migrate from salt bridge and enter into silver compartment. Electrons flow from zinc half cell to silver half cell.

Answers to Exercises

Exercise 4.8

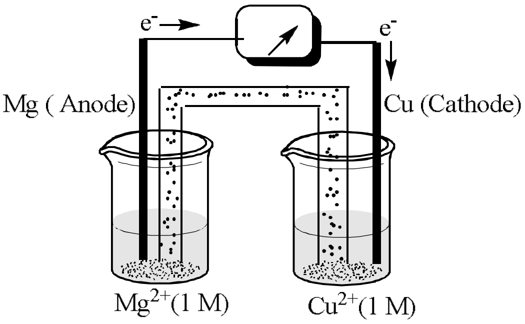
1. $Al\ |\ Al(NO_3)_3\ (1\ M)\ ||\ AgNO_3\ (1\ M)\ |\ Ag$
2. A salt bridge is used to complete a circuit or to maintain the electroneutrality of the solution.
3. a. A Galvanic cell converts chemical energy to electrical energy but in an electrolytic cell, electrical energy is used to bring about chemical change.

   b.

	Electrolytic Cell	Galvanic Cell
Cathode	Negative	Positive
Anode	Positive	Negative

Exercise 4.9

1. +0. 34 V
2. a.



b. Mg | Mg^{2+} (1M) is anode, and Cu^{2+} (1M) | Cu is cathode.

c. Electrons flow from anode (magnesium) to cathode (copper) in the external circuit.

Exercise 4.10

1. 411 kJ

Exercise 4.11

1. E_{cell} = +0.013 V. Since E_{cell} is greater than zero the reaction is spontaneous.

2.

E_{Cell}	Δ G	Cell reaction
>0	>0	Spontaneous
>0	>0	Spontaneous
=0	=0	Equilibrium

Exercise 4.12

1. A dry cell is a primary cell which is used only once and then discarded but lead storage battery is a secondary cell and can be used more than once by recharging it.

2 Cathode (reduction): $PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$

Anode (oxidation): $PbSO_4(s) + 2H_2O(l) \longrightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

Overall reaction: $2PbSO_4(s) + 2H_2O(l) \longrightarrow Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$

3. a. Ammonium chloride and zinc chloride in water

b. Anode (oxidation or Reducing agent): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
Cathode (reduction or Oxidizing agent):
$2NH_4^+(aq) + 2MnO_2(s) + 2e^- \rightarrow 2Mn_2O_3(s) + 2NH_3(aq) + H_2O(l)$

Exercise 4.13

1. Corrosion is a deterioration of a metal by redox process, which converts a metal to unwanted compound.

2. Corrosion destroys metals and returns them to their natural ores. As a result, it has negative economic impact.
3. Corrosion can be prevented by painting or coating a metal with zinc, tin or chromium.

Answers to the Review Questions

Part I Multiple Choices

1.	D	2.	B	3.	B	4.	C	5.	C
6.	B	7.	C	8.	C	9.	D	10.	C
11.	D	12.	D	13.	D	14.	B	15.	B

Part II Answers to Short Answers

16. a. Oxidation is the loss of electrons and is identified by increase in oxidation number.
    b. Reduction is the gain of electrons and is identified by decrease in oxidation number.
    c. Anode is an electrode in which oxidation takes place.
    d. Cathode is an electrode in which reduction takes place.
    e. Electrolytic cell is a cell in which electrical energy is used to bring about chemical change.
    f. Galvanic cell is a cell that converts chemical energy to electrical energy.
    g. Corrosion is the deterioration of metals by redox reactions.
17. Corrosion can be prevented by painting, galvanizing with zinc or electroplating with chromium.
18. a $4K_2CrO_4 + 4H_2O + 3S \longrightarrow 3SO_2 + 8KOH + 2Cr_2O_3$
    b $3CuO + 2NH_3 \longrightarrow 3Cu + N_2 + 3H_2O$
    c $2H_2SO_4 + 2NaBr \longrightarrow Br_2 + SO_2 + Na_2SO_4 + 2H_2O$
19. a $Mn^{2+} + BiO_3^- \longrightarrow MnO_4^- + Bi + 2H^+$
    b $4H_2O_2 + Cl_2O_7 + 2OH^- \longrightarrow 2ClO_2^- + 4O_2 + 5H_2O$
20. a. $4MnO_4^- + 5CH_3OH + 12H^+ \longrightarrow 4Mn^{2+} + 5HCOOH + 11H_2O$
    b. $3NO_2^- + Cr_2O_7^{2-} + 8H^+ \longrightarrow 2Cr^{3+} + 3NO_3^- + 4H_2O$
21. $\Delta G^o = -311$ kJ
22. Volume of hydrogen gas is 6.85 L and that of oxygen gas is 3.43 L

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UNIT 5

SOME ELEMENTS IN NATURE AND INDUSTRY

UNIT OVERVIEW Total Periods allotted: 17 periods

This unit is divided into two manageable subunits. The first subunit focuses on some elements in nature. It explains the major components of the Earth's crust, and the distribution of elements in the Earth's crust and in the atmosphere. It also discusses the carbon, nitrogen and phosphorus cycles in nature, and how human beings interfere with the natural cycles, especially to increase agricultural productivity and to produce energy.

The second subunit discusses how elements are extracted from their natural ores using different physical and chemical methods such as gravity separation, magnetic separation, froth flotation, amalgamation, leaching, chemical reduction, distillation, electrolysis and zone refining.

It also explains the extraction, properties and uses of some elements such as sodium, calcium, tin, lead, zinc, chromium and silicon. The last section of this subunit describes the preparation and uses of ammonia, nitric acid, sulphuric acid and diammonium phosphate.

This unit requires comprehensive understanding of physical and chemical principles that students have learned in chemistry at different grade levels. Therefore, they have to be encouraged to apply the physical and chemical methods that are applicable to extract elements from their ores and to convert them into practically useful materials.

Active learning methods such as pair discussion, group discussion, demonstration, presentations and other appropriate methods can be used to teach this unit.

Unit outcomes

After completing this unit, students should be able to:

- *describe occurrence and abundance of the elements in nature;*
- *explain how carbon, nitrogen and phosphorus cycles exist in nature;*
- *understand the metallurgical processes;*

- *understand the occurrence, extraction, chemical properties and uses of sodium, calcium, tin, lead, zinc and chromium;*
- *describe the occurrence, extraction, chemical properties and main uses of silicon;*
- *explain the major steps in the industrial production of ammonia, nitric acid, sulphuric acid and diammoniummonohydrogen phosphate;*
- *demonstrate scientific enquiry skills: classifying, communicating, asking questions, relating cause and effect and making generalizations.*

Main Contents

5.1Some Elements in Nature

5.2 Some Elements in Industry

5.1 SOME ELEMENTS IN NATURE

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *define the terms mineral and ore;*
- *describe the forms of occurrences of metals;*
- *discuss the distribution and relative amounts of the elements in the earth's crust;*
- *define the term fixation;*
- *discuss the carbon cycle, the nitrogen cycle and the phosphorus cycle.*

Forward Planning

Dear Colleague, before you begin this subunit, it is advisable to read contents of the subunit from the Student Text Book and other reference books available in the school library. In addition, read the Activities and suggested answers from this Teachers Guide, and plan convenient ways of forming groups.

Teaching Aids

Diagrams of carbon, nitrogen, and phosphorus cycles

Subject Matter Presentation

For this topic, it is possible to use group discussion, presentation, and question and answer as active learning methods.

You can start this subunit by allowing students to perform the Start-up Activity. Tell them to bring different pieces of metals and non-metals from their neighbourhoods before you begin this lesson, and encourage them to present to the class how they classified the elements as metals and non-metals. In their presentation, help them to describe the general properties of metals and non-metals.

In addition, students need to be encouraged to refer to books about man-made elements, and present their investigation to the class.

Direct them to come up with the following answers:

1. Metals and non-metals are identified by their properties. Metals have lustre or shiny appearance. They conduct electricity. However, non-metals do not have lustre and cannot conduct electricity. These two properties can help them to classify metals and non-metals.
2. The answer depends on the type of element selected by students. They should look for three man-made elements from the periodic table and describe the physical properties of the elements such as physical state at room temperature, density, melting point and etc.

Ask them the number of elements in the periodic table and remind them that only 92 elements are found in nature and that the others are made from them.

Help them to form groups and do **Activity 5**.1. After discussion, encourage them to share their ideas with the other groups. This activity is designed to remind them of the elements found in their body and surrounding.

Harmonize their discussions and direct them to come up with the following answers:

a. Nitrogen and oxygen are the major gaseous elements in the atmosphere.

b. Many metals such as iron, aluminium, copper, gold etc. can be listed.

c. Elements such as nitrogen and oxygen are gaseous at room temperature, but most metallic elements are solids

d. Non-metallic elements such as phosphorus, nitrogen, hydrogen, oxygen, sulphur and etc are found in our body in the form of compounds. We also use oxygen for respiration.

Ask them why elements like sodium, potassium and calcium exist in compound form, while others such as gold, platinum, and ruthenium are found free in nature. Help them to associate occurrences of elements with reactivity of the elements. Most reactive elements found at the top of the reactivity series exist in the form of compounds and the less reactive elements found at the bottom of the reactivity series exist in native state.

Encourage them to write the chemical formula and name of the common ores of sodium, calcium, magnesium, tin, lead, mercury, iron and etc. Writing the name and formula of these ores may help them to understand the extraction of some of these metals from their ores.

Then, give them **Exercise 5.1**in the Student Text as class work. Conclude this section by reiterating the main points of the lesson and giving feedback on their responses.

Recycling of Elements in Nature

For this subunit, it is possible to use brainstorming, gapped lecture, group discussion, and presentation as active learning methods

You can begin the carbon cycle by letting students do **Activity 5.2**. Encourage students to discuss the impacts of afforestation, deforestation and building of large cement factories. Except for afforestation, all the others increase global warming. Students should aware that one way of reducing global warming is afforestation. Plants use carbon dioxide to prepare glucose and thereby remove it from the atmosphere. Students should develop the habit of planting trees and conserving the environment.

You can use gapped lecture to discuss the recycling of elements in general and carbon in particular between living and nonliving things. Then, allow students to discuss in groups the diagram of the carbon cycle found in their text book, and tell them to assign group representatives and explain the carbon cycle to the class.

Give them Exercise 5.2as class work to discuss in groups and give their answers. Then end this session by providing feedback on their responses.

You may start the nitrogen cycle by letting them do Activity 5.3in groups**.** After discussion, encourage them to share their ideas with the class. The activity helps them to understand the uses of nitrogen-containing fertilisers to boost agricultural productivity. It also helps them to understand how legumes convert atmospheric nitrogen to nitrogen compounds that can be used to synthesize proteins and nucleic acids.

Even though nitrogen containing fertilisers are very important for plant growth and increase agricultural productivity, they have disadvantages. The nitrogen from fertilisers and manures are eventually converted by bacteria in the soil to nitrates. These nitrates can be leached into the groundwater or be washed out of the soil surface into streams and rivers. High nitrate levels in drinking water are considered to be dangerous to human health. They can also pollute aquatic environments by enhancing algal growth in water bodies.

Ask them about the importance of nitrogen in living organisms. It should be emphasized that the triple bond in nitrogen makes it less reactive.

Then, help students to form groups and tell them to discuss the difference between nitrogen fixation and denitrification using their text book. Ask them to identify the natural and chemical processes that convert atmospheric nitrogen into nitrogen compounds and the way nitrogen returns to the atmosphere by the decomposition of dead organisms. Encourage students to discuss the diagram of the nitrogen cycle found in the student text book with their group members. Allow them to present what they understood from the diagram to the class. Appreciate these students for their presentations, give a brief summary about the nitrogen cycle and conclude this session by giving Exercise 5.3 as homework.

You can start the phosphorus cycle by reviewing the nitrogen and carbon cycles. Raise the following questions for them to brainstorm on:

What makes the phosphorus cycle different from the carbon and nitrogen cycles?

For what purpose is phosphorus required by living organisms?

Give them ten minutes to discuss the two questions and forward their answers. Comment on their answers and provide them with a brief lecture about the phosphorus cycle for about 15 to 20 minutes (gapped lecture).

Then give them **Exercise 5.4** as class work for 15 minutes. Encourage them to reflect their answers. Conclude this section by giving feedback on their answers.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing the start-up activity and the activities from Activity 5.1 to Activity 5.3
- ✓ presenting their views after discussion,
- ✓ answering the questions in Exercise 5.1 and Exercise 5.4

By observing their performances from the record, provide them feedback to improve students' learning (formative assessment). In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

1. We are currently worried that CO_2 sources are out of balance with CO_2 sinks. If sources produce more CO_2 than sinks can remove, then CO_2 in the atmosphere

increases, possibly leading to global warming. What might happen if the reverse were true and sinks took up more CO_2 than sources?

2. Why is ocean pollution harmful to the environment?

*3. Why is deforestation harmful to the environment?

*4. How does the nitrogen cycle contribute the formation of acid rain?

Answer to Additional Questions

1. Less CO_2 would be available for plant growth or less atmospheric CO_2 might result in a slightly cooler atmosphere.
2. Ocean pollution can cause algae to die and therefore not remove carbon dioxide from the atmosphere
3. Without the plants, a lot of carbon dioxide is not removed from the atmosphere.
4. During lightning, nitrogen combines with oxygen to form nitric oxide. The nitric oxide will be oxidized to nitrogen dioxide which dissolves in rain to form nitric acid.

Answers to Exercises

Exercise 5.1

1. Whether elements exist in the form of compounds or in native state depends on their reactivity. The most reactive elements exist in the form of compounds and the least reactive elements exist in native state.
2. Oxygen, silicon, aluminium, iron, calcium, and magnesium. Oxygen is non-metal, silicon is metalloid, and the others are metals.
3. A mineral is a solid substance that contains an element. An ore is a mineral in which the element can be extracted easily and economically.
4. a. Galena b. Pyrites c. Epsomite
   d. Cinnabar e. Hematite f. Dolomite

Exercise 5.2

1. a. Green plants incorporate carbon atoms from carbon dioxide during

   Photosynthesis, but animals incorporate carbon atoms by eating plants or animals that are eating plants or other animals.

   b. Carbon returns to the atmosphere by respiration, burning of fossil fuels, and decomposition of dead organisms.

2. This is due to the fact that as temperature increases, the solubility of the gas in liquid decreases. As a result, more carbon dioxide gas will be released to the atmosphere as the temperature increases.

Exercise 5.3

1. Plants and animals need nitrogen to synthesize proteins and nucleic acids (RNA and DNA).
2. Because the nitrogen–nitrogen triple bond has high bond dissociation energy (945 kJ/mol).
3. Nitrogen fixation is the conversion of atmospheric nitrogen into nitrogen compounds, but denitrification is the reverse process.
4. Naturally, nitrogen is fixed by lighting and bacteria in the root nodules of leguminous plants. Industrially, nitrogen is mainly fixed by the Haber process.

Exercise 5.4

1. Phosphorus cannot be readily washed out of the soil, but is bound to soil particles and moves together with them. Phosphorus can therefore be washed into surface waters together with the soil that is being eroded. The phosphorus is not considered to be dangerous, but it stimulates the growth of algae in slow moving water. These algae eventually die and decompose, removing the oxygen from the water causing fish kills. This process is called eutrophication.
2. In the case of carbon and nitrogen cycles, we can easily notice that both exist in gaseous form, but phosphorus does not exist in gaseous form. In addition, the phosphorus cycle is very slow. Phosphorus remains in rocks sediments for millions of years.
3. Cells use phosphorus to make DNA, RNA and ATP.

5.2 SOME ELEMENTS IN INDUSTRY

Periods Allotted: 12 periods

Competencies

At the end of this section, students will be able to:

- *define metallurgy;*
- *explain the major steps in metallurgical processes;*
- *describe the manufacture of sodium by Down's cell*;
- *explain the chemical properties of sodium;*
- *describe the uses of sodium;*
- *describe the manufacture of calcium;*
- *explain the chemical properties of calcium;*

- *describe the uses of calcium;*
- *describe the manufacture of tin;*
- *explain the chemical properties of tin;*
- *describe the uses of tin;*
- *describe the manufacture of lead;*
- *explain the chemical properties of lead;*
- *describe the uses of lead;*
- *describe the manufacture of zinc;*
- *explain the chemical properties of zinc;*
- *describe the uses of zinc;*
- *describe the manufacture of chromium;*
- *explain the chemical properties of chromium;*
- *describe the uses of chromium;*
- *describe the production of silicon;*
- *explain the chemical properties of silicon;*
- *describe the uses of silicon;*
- *explain the steps in Haber process in the industrial production of ammonia and its uses;*
- *explain the steps in Ostwald's process in the industrial production of nitric acid and its uses;*
- *explain the steps in Contact process in the industrial production of sulphuric acid and its uses;*
- *explain the steps in the industrial production of diammoniummonohydrogenphosphate and its uses.*

Foreword Planning

Students need your help in giving directions and clarifying vague concepts for them. They also need to know the answers of the activities and exercises. Refer to the Student Text Book and the answers of the Activities and Exercises in this Teachers Guide before you start this subunit. It is also advisable to refer to the reference books available in the school library.

Subject Matter Presentation

Metallurgy

For this subunit, it is possible to use gapped lecture, brainstorming, experiment, group discussion, and presentation as active learning methods.

Begin this subunit by encouraging students to do **Activity 5.4** in groups. This activity is designed to help them recap the application of electrolysis in the extraction of elements. Help them by facilitating group formation and discussion. Help them to look into the major steps to be carried out to extract metals from their natural ores. Special emphasis should be given to the physical and chemical methods that are applicable for the extraction of metals.

Direct them to come up with the following answers:

1. The major steps that are carried out to get a pure metal from its ore are: concentrating the ore, pre-treatment, extracting the metal and refining.
2. The reactivity of a metal determines the use of either chemical or electrolytic reduction method of extraction.
3. The most reactive metals such as alkali and alkaline earth metals can be extracted from their ores by electrolytic reduction because they cannot be reduced by the common reducing agents like coke. The less reactive metals like zinc, iron, lead, etc. can be reduced by carbon.

As indicated in question 3, the most reactive metals are extracted from their ores by electrolysis. Metals which are less reactive can be obtained from their ores by chemical reduction.

Then, buttress the discussion by raising the following question.

What are the major steps carried out to obtain a pure metal?

Their answers to this question may help you to identify students' understanding of the lesson. If students do not have good knowledge of the topic, tell them that the major steps carried out to obtain a pure metal are concentrating the ore, pre-treatment, production of the metal, and refining and alloying. Then, give them a brief explanation about the methods of concentrating an ore such as gravity separation, magnetic separation, froth flotation, amalgamation, and leaching.

Then let them do Activity 5.5in groups. This activity is aimed at reminding them of the applications of amalgams in their daily life. After discussion, help them to share their ideas with the rest of the class. Then comment on their answers and harmonize their discussion by giving the following answers:

1. Amalgam is an alloy of mercury with another metal. Except for iron and platinum, all metals dissolve in mercury.
2. Non-metals do not form amalgams. This is because metal atoms in metals are held by metallic bonding and non-metal atoms are held by weak intermolecular forces such as Van der Waals forces. Hence, the force of attraction between mercury and non-metals are different, and they cannot form amalgams.
3. Amalgams are commonly used by dentists to fill cavities in teeth.

If students understand the topic well, let them discuss in groups the methods for concentrating an ore. You can facilitate their discussion by clarifying vague concepts for them.

Before you begin the lesson on pre-treatment, review the methods of concentrating ores which they learned in the previous period. You can start this section by giving a brief explanation of the pre-treatment of ores, chemical and electrolytic reductions, and methods of purifying metals such as distillation, electrolysis, and zone refining.

For pre-treatment of ores, you can use the roasting of Pb Sand the conversion of Mg^{2+} ions to $MgCl_2$ as examples.

Metals are produced either using chemical reduction or electrolytic reduction. Ask students to identify the property that determines the application of chemical or electrolytic reduction.

We use chemical reduction if the metal to be extracted is less electropositive like Zn, Fe, Sn, and Cr. Electrolytic reduction is used to extract more electropositive metals such as Na, Ca, Mg and Al.

For refining of metals, have them focus on the purification of mercury, magnesium and zinc by distillation, the purification of copper by electrolysis and the general principles of purification by zone refining.

Let them do Activity 5.6 in groups and encourage them to present their answers to the class. This activity lets students investigate elements that are found in sea water and rivers. It also helps them to identify the major elements that can be economically extracted from sea water. Finally, comment on their answers and direct them to come up with the following answers:

1. The main elements found in sea water and rivers are: sodium, potassium, calcium, magnesium, etc.
2. Yes.
3. The alkali and alkaline earth metals.

Sodium

In lower grades, students learned the chemistry of alkali metals. So, help them to form groups and do Activity 5.7.This activity is aimed at inspiring them list down the main ores of sodium which are used for the extraction of sodium metal and its chemical properties. In addition, it reminds students of the major ores of sodium that are produced in Ethiopia. After discussion, help them to share their ideas with the rest of the class. Harmonize their discussions and guide them to come up with the following answers:

1. Borax ($Na_2B_4O_7.10H_2O$), Trona ($Na_2CO_3.NaHCO_3.2H_2O$), Rock salt (NaCl)
2. Trona is found in Lake Abiyata, and Rock salt is found Lake Afdera.
3. Afar region.
4. Because sodium reacts with both oxygen and water.

5. The answers are given in Activity 4.3 (c).

In addition, tell them to discuss the extraction, chemical properties and uses of sodium metal in their groups. Facilitate the group discussions by clarifying vague concepts (if any). Encourage them to share what they discussed with the other groups or the class.

Conclude this section by giving them Exercise 5.7 as homework.

Calcium

Similarly, students learned the chemistry of alkaline earth metals in lower grades. So, they are familiar with the chemistry of calcium.

Therefore, to make this lesson more attractive, start it by raising the following questions for them to brainstorm on as much as they can;

- ✓ List down the main ores of calcium.
- ✓ Write the reactions that take place at the cathode and anode in the electrolysis of molten $CaCl_2$.
- ✓ Write a chemical equation that shows the reaction between calcium and:

  a. H_2O b. F_2 c. H_2

Then let them form groups and discuss the occurrence, extraction, chemical properties and uses of calcium. They can get the answers for the above questions during their discussion. Let some group representatives present what they discussed to the class and comment on their presentations.

Conclude this section by recapitulating the main concepts and by giving them

Exercise 5.8 as homework.

Tin

Students did not encounter the chemistry of tin in lower grades. Therefore, it is recommended to give them a brief explanation about the occurrence, chemical properties and uses of tin. Give them Activity 5.8ahead of time to refer to chemistry books about the extraction of tin from cassterite and present their findings to the class. This Activity is intended to acquaint students with the use of reference books or scientific literatures from the library or any sources.

Tin is extracted from SnO_2 by using carbon as a reducing agent, using the following chemical reaction: $SnO_2 + C \rightarrow Sn + CO_2$.

Ask them why electrolytic reduction is not used for its extraction.

Then let them discuss in groups and answer the questions in **Exercise 5.9**. Ask them to assign group representatives to present their answers to the class. Finally, give feedback and credit to the presenters.

Dear colleague, giving credit for active learners may motivate other students to participate in group discussions and presentations.

Lead

As you did for tin, introduce lead by giving a brief discussion (gapped lecture) about its occurrence, the two methods of extracting it, (roasting followed by reduction with carbon monoxide and partial oxidation of galena followed by self reduction), and its chemical properties and uses.

Then allow them to answer the questions in Exercise 5.10 and ask them to present their answers to the class. Give credits to the presenters and comment on their answers.

Zinc

You can start the lesson for zinc by raising the following brainstorming questions;

- ✓ What is the most important ore of zinc?
- ✓ How is zinc extracted from its ore?
- ✓ Write the products if zinc reacts with dilute HCl and oxygen.
- ✓ What are the major uses of zinc?

Let them form groups and discuss the occurrence, extraction, chemical properties and uses of zinc. They will get the answers to the above questions during their discussions.

Ask them to assign group representatives for class presentation. Encourage group presenters and end this session by giving them Exercise 5.11 as homework.

In Experiment 5.1 students should observe some chemical reactions by doing the experiment in groups or by demonstration. It is believed that students learn more by doing than listening. So, allow them to investigate the reactions and colour changes of indicators in acids and bases. In addition, students should be encouraged to write chemical equations for the reactions of dilute HCl with Sn, Pb and Zn. Then ask them to write a laboratory report.

The answers to the questions in the Results and Discussion section of the experiment are as follows:

1. Hydrogen gas. It gives a popping sound when it burns with oxygen.
2. a. $H_2 + O_2 \rightarrow H_2O$

   b. $Zn + HCl \rightarrow ZnCl_2 + H_2O$

Chromium

In the extraction of chromium, rather complex chemical reactions are involved. So we advise you to use gapped lecture (giving a brief lecture about 15 to 20 minutes) to discuss the occurrence, extraction and uses of chromium. Then allow students to form groups and do **Exercise 5.12.** Encourage them to present their answers to the class. Give feedback on their responses.

Silicon

You can begin the lesson of silicon by asking students and letting them brainstorm on the following questions:

- ✓ What is the second most abundant element in the Earth's crust?
- ✓ How is pure silicon obtained from silica?
- ✓ Which acid attacks glass?
- ✓ What are silicones?
- ✓ What are the main uses of silicon?

Allow students to do Activity 5.9.This activity emphasizes the reactivity of metals. Help them to form groups and answer the above questions by reading and discussing the chemistry of silicon in their text book. Facilitate the group discussions by clarifying unclear concepts. Harmonize their discussion and direct them to come up with the following answers:

Elements like tin, lead, zinc, chromium, and silicon are less reactive elements and can be reduced by carbon, but sodium and calcium are very reactive elements and cannot be reduced by carbon.

After they come up with consensus, let them do the same thing for Activity 5.10. This activity is aimed at helping students to investigate elements that react with air and water. Then encourage them to assign group representatives and present their answers to the class. Comment on their answers and direct them to come up with the following answers:

1. Elements such as silver, gold, platinum, palladium, ruthenium and osmium do not react with water and air.
2. Metals such as alkali and alkaline earth metals react vigorously with water
3. Since alkali and alkaline earth metals react with water and air, they should be kept under kerosene oil.

Ammonia

Help them to form groups and discuss the questions in Activity 5.11. This activity reminds them to apply the factors that affect chemical equilibrium in the synthesis of ammonia by Haber process. Facilitate their group discussions. After the discussion, encourage them to share their ideas with the class. Harmonize their discussion and help them to come up with the answers which are found in the "Note for the Teacher" section.

Nitric Acid

You can start the lesson on nitric acid by using gapped lecture. Give them a brief lecture for 15 minutes about the physical properties, industrial preparation and uses of nitric acid. Then let students form groups and give them **Exercise 5.14** as class work to be presented to the class. Appreciate and give credits to the presenters. Provide them with feedbacks on their answers and conclude the lesson.

Sulphuric Acid

You can use group discussion or any other appropriate active learning methods to teach sulphuric acid.

You can begin this lesson by raising the following questions for brainstorming or as guiding questions:

- ✓ How is the consumption of sulphuric acid related to the development of a country?
- ✓ What are physical properties of sulphuric acid?
- ✓ Show the preparation of sulphuric acid by the contact process, using chemical equations.
- ✓ What are the main uses of sulphuric acid?

The answers of the above questions are already found in the student text. Let them form groups and discuss and answer the questions together by reading their text books. Then encourage them to presenters their answers to the class by assigning group representatives. Appreciate the presents and give them credits to motivate other students. End this session by giving them **Exercise 5.15** as homework.

Diammonium Monohydrogen Phosphate, DAP

This is the last section of the unit. First help them to form groups and discuss the questions in Activity 5.12.After discussion, encourage them to share their answers with the class. Facilitate their presentation by clarifying vague concepts for them. Harmonize their discussions and guide them to come up with the following answers:

1.
   a. Nitrogen deficiency is recognized by the yellowing of older leaves, slowing or stopping of growth. Leaves may drop sooner than expected. Excess nitrogen is recognized by extremely fast growth, resulting in long, spindly, weak shoots with dark green leaves.
   b. Phosphorus deficiency is recognized by dull green leaves and purplish stems. The plant is generally unhealthy, sometimes yellowing. Lack of blooming with lush green foliage may also indicate a lack of phosphorus.
   c. Potassium deficiency is hard to diagnose by symptom, but plants are generally sickly, with small fruit, yellowing from the older leaves upwards, and have sickly blooms.
2. The main environmental problems associated with fertilizers are contamination of water with nitrates and phosphates. The nitrogen from fertilizers and manures are

eventually converted by bacteria in the soil to nitrates. These nitrates can be leached into the groundwater or be washed out of the soil surface into streams and rivers. High nitrate levels in drinking water are considered to be dangerous to human health. The phosphorus itself is not considered to be dangerous, but it stimulates the growth of algae in slow moving water. These algae eventually die and decompose. As they decompose, they deplete oxygen from water bodies. Consequently, aquatic organisms will die due to lack of sufficient oxygen.

In their Groups and ask the following questions for the students to answer by reading their text books.

- ✓ What are the physical properties of DAP?
- ✓ Show the preparation of DAP, using chemical equations.
- ✓ What are the uses of DAP?

Let them discuss together, answer the above questions and do **Exercise 5.16**. Then encourage them to present their answers to the class by assigning group representatives. Appreciate the presenters and give them credits. End this session and unit by giving feedback on their answers and by asking them to submit the answers to the review questions.

Note for the Teacher

Ammonia is one of the world's most important industrial chemicals. It is a colourless gas with a pungent odour that condenses to a colourless liquid at –33°C. Gaseous ammonia is very soluble in water because ammonia molecules are polar and can form hydrogen bonds with water molecules.

Ammonia is prepared industrially from hydrogen and nitrogen by the Haber process.

$$\underbrace{N_2(g) + 3H_2(g)}_{\textit{4 volumes}} \xrightleftharpoons{380\text{-}450^{\circ}C,\ Fe} \underbrace{2NH_3(g)}_{\textit{2 volumes}};\qquad \Delta H^{\circ} = -92.2\ \text{kJ/mol}$$

The reaction is reversible, and *Le Chatelier's* principle suggests that high pressure and low temperature are required to drive the reaction to the right, and thus form NH_3. A low temperature gives a higher percentage of conversion to NH_3, but the reaction is slow in reaching equilibrium, and a catalyst is required.

In practice, the conditions used are 200 atm, 380 – 450°C temperature and a catalyst of finely-divided iron. It is more economical to use high temperature, which causes equilibrium to be reached much faster, even though this condition gives a lower percentage of conversion. At a temperature of about 400°C, a 15% conversion is obtained with a single pass over the catalyst. The gas mixture is cooled to condense liquid ammonia, and the unreacted mixture of N_2 and H_2 gases is recycled.

Ammonia can be prepared in the laboratory by heating ammonium chloride with sodium hydroxide:

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O\ (l) + NH_3(g)$$

Most of the ammonia produced is used to make fertilisers. It is also used for making HNO_3, which can be used to make NH_4NO_3 (*fertiliser*), or explosives such as nitroglycerine, nitrocellulose and trinitrotoluene (TNT). Ammonia has been used as the cooling liquid in refrigerators. It has a very high heat of vaporization, and convenient boiling and freezing points.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing during the activities from Activity 5.4 to Activity 5.12,
- ✓ presenting their views after discussion,
- ✓ answering the questions from Exercise 5.5 to Exercise 5.16

In addition, you can assess their performance by letting them reflect their laboratory report for Experiment 5.1. By observing their performances from the record, provide them with feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit. You can also use the review questions of this unit as part of summative assessment.

Additional Questions

1. How does roasting differ from calcination?
2. Why are aluminium containers used to transport nitric acid?
3. Carbonates and sulphide ores are usually converted into oxide ores. Why?

*4. The following questions are from the electrolysis of aluminium oxide.

   a. Why does the anode need to be replaced regularly?

   b. Write an equation to show how aluminium ions are changed in to aluminium metal.

   c. Why can't the blast furnace be used to make aluminium?

d. Why is cryolite added to the aluminium ore during electrolysis?

Answers to Additional Questions

1. Roasting is heating ore in the presence of air to oxidize it. Calcination is heating of the ore, in the absence of air, to decompose it or drive out volatile matter. Roasting occurs at higher temperatures, higher than the melting point of the ore. Calcinations occurs at temperatures lower than the melting point of the ore
2. Aluminium containers are used to transport nitric acid because concentrated HNO_3 makes aluminium passive by forming a thin strong protective oxide layer. This protective layer prevents further reaction with the acid.
3. It is easier to reduce oxide as compared to carbonate and sulphide
4. a. The carbon/graphite anode reacts with the oxygen made to form carbon dioxide gas.
   b. $Al^{3+}(l) + 3e^- \rightarrow Al(l)$
   c. Aluminium is more reactive than carbon (not too reactive).
   d. To lower the melting point of the aluminium oxide (so less energy required)and to increase electrical conductivity.

Answers for Exercises

Exercise 5.5

1. Metallurgy is the science and technology of extracting metals from their natural sources and preparing them for practical use.
2. a. A gravity separation is a method of separating an ore from its gangue based on density differences between the ore and the gangue. For example, gold is separated from its gangue by gravity separation.
   b. Magnetic separation is applied to separate ferromagnetic substances from their impurities. For example, magnetite is separated from its gangue using a magnet.
   c. Froth Flotation is applied based on their solubility difference between mineral and its gangue with oil and detergent. For example, sulphide ores are concentrated by flotation.
   d. Amalgamation is employed to extract metals from their ores. For example, gold and silver are extracted using this method.
   e. Leaching is a chemical method of concentrating ores. For example, gold and aluminium are concentrated by leaching.

Exercise 5.6

1. The first method is concentrating an ore using physical and chemical methods. Once the ore of a metal is concentrated, it will be reduced by using chemical or

electrolytic reduction; then if the metal is impure, it will be purified by distillation, electrolysis or zone refining.

2. Pretreatment is a method of converting an ore into a compound that is easy to reduce.

   For example, galena (PbS) is converted to lead oxide by roasting with air. The lead oxide is easily reduced to give lead metal by carbon.

3. The chemical reduction method is applied for less electropositive metals. For example,

$$\underset{\substack{\uparrow \\ \text{less} \\ \text{electropositive}}}{Cr_2O_3(s)} + \underset{\substack{\nwarrow \\ \text{more} \\ \text{electropositive}}}{Al(s)} \longrightarrow 2Cr(l) + Al_2O_3(s)$$

   The electrolytic reduction method is used to extract the more electropositive metals like sodium, calcium and aluminium.

4. In zone refining, a metal rod containing impurities is usually heated with an electrical heating coil. As the metal rod moves slowly, the molten metal cools and forms solid, but the impurities still remain in liquid state and move together with the electrical heating coil.
5. Lead and zinc are oxidized because they have lower reduction potential than copper, but silver and gold have higher reduction potential than copper and remain as anode sludge.

Exercise 5.7

1. a. $Na_2B_4O_7.10H_2O$ b. $Na_2CO_3.NaHCO_3.2H_2O$ c. NaCl
2. To decrease the melting point of NaCl from 800°C to 600°C.
3. Because sodium reacts with water and releases heat energy.
4. a. $2Na(s) + 2Cl_2(g) \rightarrow 2NaCl(s)$

   b. $2Na(s) + Br_2(g) \rightarrow 2NaBr(s)$

   c. $2Na(s) + I_2(g) \rightarrow 2NaI(s)$
5. Sodium is used to make anti-knock, to reduce metals like zinc from its compounds, and make sodium vapour lamp for street lights.

Exercise 5.8

1. Limestone ($CaCO_3$), dolomite ($CaCO_3 . MgCO_3$), and gypsum ($CaSO_4.2H_2O$).
2. Step 1. $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

   Step 2. $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

   Step 3. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

3. a. $Ca(s) + Cl_2(g) \longrightarrow CaCl_2(s)$
   b. $Ca(s) + Br_2(l) \longrightarrow CaBr_2(s)$
   c. $Ca(s) + I_2(s) \longrightarrow CaI_2(s)$
   d. $3Ca(s) + N_2(g) \xrightarrow{\text{high temperature}} Ca_3N_2(s)$
4. Calcium is an essential element in living matter. It is the major component of bones and teeth. A characteristic function of Ca^{2+} ions in living systems is in the activation of a variety of metabolic processes. For example, calcium plays a vital role in heart action, blood clotting, muscle contraction, and nerve impulse transmission.

Exercise 5.9

1. $SnO_2(s) + 2C(s) \xrightarrow{\Delta} Sn(l) + 2CO(q)$
2. (i) $Sn(s) + H_2O(g) \xrightarrow{\Delta} SnO_2(s) + H_2(g)$

   (ii) $Sn(s) + HCl(aq) \xrightarrow{\Delta} SnCl_2(aq) + H_2(g)$
3. $Sn(s) + 2Cl_2(g) \xrightarrow{\text{cold}} SnCl_4(s)$

   $Sn(s) + 2Br_2(l) \xrightarrow{\text{cold}} SnBr_4(s)$

   $Sn(s) + 2I_2(s) \xrightarrow{\text{heat}} SnI_4(s)$
4. Tin is mainly used for the electroplating of iron or steel and making alloys such as bronze and solder.

Exercise 5.10

1. i. Roasting in air followed by reduction with coke.

   $2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + SO_2(g)$

   $2PbO(s) + C(s) \xrightarrow{\Delta} 2Pb(l) + CO_2(g)$

   ii. Partial oxidation followed by self-reduction

   $3PbS(s) + O_2 \xrightarrow[\text{air}]{\text{heat in}} PbS(s) + 2PbO\ (s) \xrightarrow[\text{absence of air}]{\text{heat in}} 3Pb(l) + SO_2(g)$
2. a. No reaction because the surface of lead is covered by lead oxide.
   b. $Pb(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2(g)$
   c.
      (i) With dilute nitric acid : $3Pb(s) + 8HNO_3(aq) \rightarrow 3Pb(NO_3)_2(aq) + 2NO(g) + 4H_2O$

      (ii) With concentrated nitric acid : $Pb(s) + 4HNO_3(aq) \rightarrow Pb(NO_3)_2(ag) + 2NO_2(g) + 2H_2O\ (l)$
3. Lead is used to make car batteries, paints, pipes and alloys.

Exercise 5.11

1. i. Roasting of ZnS followed by reduction with carbon monoxide

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2$$

$$ZnO(s) + CO(g) \rightarrow Zn(g) + CO_2(g)$$

ii. Heating at lower temperature to form ZnO and $ZnSO_4$. Then, addition of H_2SO_4 followed by electrolysis.

$$ZnS(s) + O_2(g) \rightarrow ZnO(s) + ZnSO_4(s)$$

Cathode (reduction): $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

2. Zinc is amphoteric because it reacts with both acids and bases.
3. Zinc is used for galvanizing iron and making alloys.

Exercise 5.12

1. i. $FeCr_2O_4(s) + 4C(s) \xrightarrow[furnance]{electric} Fe(l) + 2Cr(l) + 4CO(g)$

ii. $Cr_2O_3(s) + Al(s) \xrightarrow{\Delta} 2Cr(s) + Al_2O_3(s)$

2. **Step 1** $FeCr_2O_4(s) + Na_2CO_3(s) + O_2(g) \rightarrow Na_2CrO_4(s) + Fe_2O_3(s) + CO_2(g)$

Step 2 $2Na_2CrO_4(s) + 2H^+(aq) \rightarrow Na_2Cr_2O_7(s) + H_2O(l) + 2\overset{+}{Na}(aq)$

Step 3 $Na_2Cr_2O_7(s) + 2C(s) \rightarrow Cr_2O_3(s) + Na_2CO_3(S) + CO(g)$

Step 4 $Cr_2O_3(s) + 2Al(s) \rightarrow 2Cr(s) + Al_2O_3(s)$

Exercise 5.13

1. First impure silicon is obtained by the reduction of silica with carbon

$$SiO_2(s) + 2C(s) \rightarrow Si(l) + 2CO(g)$$

Then it is allowed to react with chlorine.

$$Si(s) + 2Cl_2(g) \rightarrow SiCl_4(g)$$

Finally, the silicon tetrachloride will be reduced by adding magnesium metal.

$$SiCl_4(g) + 2Mg(s) \rightarrow 2MgCl_2(s) + Si(s)$$

2. Silicon is mainly used in the construction of transistors and microprocessors. Silicone polymers are used in making lubricants, car polish, lipstick and other materials.

Exercise 5.14

1. Nitric acid is a colourless liquid, but it turns to brown when it is exposed to light, due to the formation of NO_2. It is a strong acid and strong oxidizing acid.
2. **Step 1:** Oxidation of ammonia

$$4NH_3(g) + 5O_2 \xrightarrow[Pt]{850^oC} 4NO(g) + 6H_2O(g)$$

Step 2: Conversion of NO to NO_2

$$2NO(g) + O_2(g) \xrightarrow{cool} 2NO_2(g)$$

Step 3: Dissolution of NO_2 with water

$$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO$$

NO is recycled.

3. It is mainly used for preparing fertilizers, explosives, and plastics.

Exercise 5.15

1. Sulphuric acid is prepared by contact process.

Step 1: Burning of Sulphur

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Step 2: Conversion of SO_2 to SO_3

$$2SO_2(g) + O_2(g) \xrightarrow[400^oC]{V_2O_5} 2SO_3(g)$$

Step 3: Passing SO_3 to sulphuric acid

$$SO_3(g) + H_2SO_4(aq) \rightarrow H_2S_2O_7(l)$$

Step 4: Addition of water to oleum

$$H_2S_2O_7(l) + H_2O(l) \rightarrow 2H_2SO_4(aq)$$

2. The major uses of sulphuric acid are for making fertilizers, detergents, plastics & paints.

Exercise 5.16

1. **Step 1:** Reaction of ammonia with H_3PO_4

$$3NH_3(g) + 2H_3PO_4(l) \rightarrow (NH_4)_2HPO_4(s) + NH_4H_2PO_4(s)$$

Step 2; Recycling monoammonium phosphate

$$NH_3(g) + NH_4H_2PO_4(s) \rightarrow (NH_4)_2HPO_4(s)$$

2. DAP is a white solid which is completely soluble in water. Since it contains both nitrogen and phosphorus, it is mainly used as a fertilizer.

3. Molar mass of $(NH_4)_2HPO_4$ = 132 g/mol

$$P\% = \frac{31}{132} \times 100\% = 23.48\%$$

$$N\% = \frac{28}{132} \times 100\% = 21.21\%$$

Answers to review questions

Part IMultiple Choice

1.	B	b.	D	3.	D	4.	B	5.	A
6.	A	7.	C	8.	D	9.	D	10.	B
11.	C	12.	C	13.	D	14.	A	15.	B
16.	A	17.	A	18.	D				

Part II Answers to Short Answers

19. a. Mineral is a solid substance which occurs naturally in the earth's crust

    b. Ore is a mineral from which the element can be extracted easily and economically.

    c. Metallurgy is the science and technology of extracting metals from their natural sources and preparing them for practical use.

20. Oxygen (O), silicon (Si), aluminium (Al), iron (Fe), calcium (Ca), magnesium (Mg), sodium (Na), potassium (K), titanium (Ti), hydrogen (H), phosphorous (P), and manganese (Mn).

21. Use the diagrams of carbon, nitrogen and phosphorus cycles indicated in the student text book.

22. **Step1:** Mining and concentrating the ore.

    Step 2: Pre-treatment of the ore.

    Step 3: Production and refining of the element.

23. a. Sodium is a very reactive metal and mainly occurs in the form of compounds. The main ores of sodium are rock salt (NaCl), borax ($Na_2B_4O_7 .10H_2O$),Trona ($Na_2CO_3.NaHCO_3 . 2H_2O$).

    b. Calcium is the 5th most abundant element in the Earth's crust. The main ores of calcium are: limestone, chalk, and marble ($CaCO_3$), dolomite ($CaCO_3.MgCO_3$), gypsum ($CaSO_4.2H_2O$), and fluorite (CaF_2).

c. Tin exists in the form of compounds. The most important ore of tin is cassiterite, SnO_2.

d. Lead also exits in the form of compounds. It is mainly extracted from *galena*, PbS.

e. Chromium is the twenty–first most abundant element by weight in the Earth's crust. The most important ore of chromium is chromite ($FeCr_2O_4$).

f. Silicon is the second most abundant element, after oxygen in the Earth's crust. It is mainly extracted from silica (SiO_2).

24. a. Sodium is extracted by the electrolysis of molten sodium chloride.

Anode (oxidation): $2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$

Cathode (reduction): $2Na^+(l) + 2e^- \longrightarrow 2Na(l)$

Overall reaction: $2Cl^-(l) + 2Na^+(l) \longrightarrow 2Na(l) + 2Cl_2(g)$

b. Calcium is prepared by the electrolysis of molten calcium chloride ($CaCl_2$).

Anode (oxidation): $2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$

Cathode (reduction): $Ca^{2+}(l) + 2e^- \longrightarrow Ca(l)$

Overall reaction: $Ca^{2+}(l) + 2Cl^-(l) \longrightarrow Ca(l) + Cl_2(g)$

c. Tin is extracted from SnO_2using carbon as a reducing agent.

$$SnO_2(s) + 2C(s) \longrightarrow Sn(l) + 2CO(g)$$

d. Lead is extracted by using following two methods:

i. Roast in air to give PbO, and then reduce with coke or CO in a furnace.

$$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + SO_2(g)$$

$$2PbO(s) + C(s) \xrightarrow{\Delta} 2Pb(l) + CO_2(g)$$

ii. PbS is partially oxidized by heating and blowing air through it. After some time the air is turned off and heating is continued, and the mixture undergoes self – reduction.

$$3PbS(s) + O_2 \xrightarrow[\text{air}]{\text{heat in}} PbS(s) + 2PbO(s) \xrightarrow[\text{absence of air}]{\text{heat in}} 3Pb(l) + SO_2(g)$$

e. Chromium is produced in two forms which are ferrochrome and pure chromium metal. Ferrochrome is mainly an alloy of chromium and iron. It is prepared by reducing chromite with carbon.

$$FeCr_2O_4(s) + 4C(s) \xrightarrow[\text{furnance}]{\text{electric}} \underbrace{Fe + 2Cr}_{\textit{ferrochrome}} + 4CO$$

The following steps should be followed to get pure chromium.

$$FeCr_2O_4(s) + 16NaOH(s) + 7O_2(g) \xrightarrow{1100^\circ c} 8Na_2Cr_2O_4(s) + 2Fe_2O_3(s) + 8H_2O(l)$$

$$2Na_2CrO_4(aq) + 2HCl(aq) \longrightarrow Na_2Cr_2O_7(s) + 2NaCl(aq) + H_2O(l)$$

$$Na_2Cr_2O_7(s) + 2C(s) \longrightarrow Cr_2O_3\ (s) + Na_2CO_3(s) + CO(g)$$

$$Cr_2O_3(s) + Al(s) \longrightarrow 2Cr(s) + Al_2O_3(s)$$

f. Silicon is prepared by heating silica with coke to approximately 3000°C in an electric furnace:

$$SiO_2(s) + 2C(s) \longrightarrow Si(l) + 2CO(g)$$

Pure silicon can be obtained by the following two reactions:

$$Si(s) + 2Cl_2(g) \longrightarrow SiCl_4(l)$$

$$SiCl_4(g) + 2Mg(s) \longrightarrow 2MgCl_2(s) + Si(s)$$

25. a. Sodium is a very reactive element. It reacts with oxygen, water, hydrogen, and halogens. Because of this reason, it should be kept under kerosene or paraffin oil.

b. Calcium also reacts with oxygen, water, hydrogen, nitrogen and halogens.

c. Tin reacts with hot water, dilute and concentrated acids, and halogens.

d. lead is not affected by water because its surface is covered by PbO.

e. Chromium is a very active metal and its surface is covered by thin oxide film. As a result it does not dissolve in oxidizing acids like concentrated sulfuric and nitric acids. However, it reacts with simple acids like HCl.

f. Silicon is very less reactive. Silica or glass is attacked by HF. Silicon can also react with methyl chloride and is used to produce silicones.

26 .a. The largest use of sodium is to make a Na/Pb alloy needed to make $PbEt_4$

and $PbMe_4$. Another important use is to reduce $TiCl_4$ and $ZnCl_4$ to the metals. Liquid sodium metal is used as a coolant in nuclear reactors. It is also used to make Na_2O_2 and NaH, and for street lights.

b. Calcium is an essential element in living matter. It is the major component of bones and teeth.

c. The main uses of tin are electroplating steel to make tin-plate, and making different alloys.

d. The major use of lead is to make lead/acid storage batteries. It is also used in the manufacture of $PbEt_4$ as an additive for petrol, in paints and pigments. In addition, lead is used for the production of lead sheets, lead pipes and solder.

e. Nichrome wire, commonly used in chemistry laboratories for flame test, is an alloy of nickel and chromium. Ferrochrome, an iron–chromium alloy, is used in the production of stainless steel. Sodium and potassium chromates and dichromates are used as pigments and corrosion inhibitors in heating and air-conditioning systems.

f. Silicon is a semiconductor and is used in the construction of transistors and microprocessors. Quartz crystals are used to control the frequency of radio and television transmissions. Silicon polymers are used in making lubricants, lipstick, car polish and other materials.

27. a. Preparation of ammonia by Haber process

$$N_2(g) + 3H_2(g) \xrightleftharpoons{380\text{-}450^{o}C,\ Fe} 2NH_3(g) \quad \Delta H^{o} = -92.2\ kJ/mol$$

b. Preparation of nitric acid by Ostwald process

Step 1: Oxidation of ammonia.

$$4NH_3(g) + 5O_2(g) \xrightarrow[Pt]{850\ ^{o}C,\ 5\ atm} 4NO(g) + 6H_2O(g)$$

Step 2: Conversion of NO to NO_2.

$$2NO(g) + O_2\ (g) \xrightarrow{cool} 2NO_2(g)$$

Step 3: Reaction of NO_2 gas with water.

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + \underset{\textit{recycled}}{NO(g)}$$

The nitric oxide (NO) is recycled in step 2:

c. Preparation of sulphuric acid by contact process.

Step 1: Burning sulphur in air

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

Step 2: Converting SO_2 to SO_3.

$$2SO_2(g) + O_2(g) \xrightarrow[400\ ^{o}C]{V_2O_5\ \text{catalyst}} 2SO_3(g)$$

Step 3: Passing SO_3 into H_2SO_4.

$$SO_3(g) + H_2SO_4(l) \longrightarrow H_2S_2O_7(l)$$

Pyrosulphuric acid
(oleum or fuming sulphuric acid)

Step 4: Addition of water to pyrosulphuric acid.

$$H_2S_2O_7(l) + H_2O(l) \longrightarrow 2H_2SO_4(aq)$$

d. Preparation of DAP

Step 1: Anhydrous ammonia reacts with phosphoric acid to form monoammonium phosphate and diammonium phosphate

$$3NH_3(g) + 2H_3PO_4(l) \longrightarrow NH_4H_2PO_4(s) + (NH_4)_2HPO_4(s)$$

Step 2: Recycling monoammonium phosphate for further reaction with anhydrous ammonia.

$$NH_4H_2PO_4(s) + NH_3(g) \longrightarrow (NH_4)_2HPO_4\ (s)$$

28. a. Major uses of Ammonia:

Most of the ammonia produced is used to make fertilizers. It is also used for making HNO_3, which can be used to make NH_4NO_3 (fertilizer), or explosives such as nitroglycerine, nitrocellulose and trinitrotoluene (TNT).

b. Major uses of Nitric acid:

The largest amount of nitric acid is used to synthesize ammonium nitrate, a water soluble fertilizer. Large quantities are also used to make plastics, drugs, and explosives such as trinitrotoluene (TNT) and nitroglycerine.

c. Major uses of Sulphuric acid

Sulphuric acid is used in the production of fertilizers, detergents, plastics and paints.

It is an oxidizing agent and a good dehydrating agent. It is also used as electrolyte in car batteries.

d. Major uses of DAP

DAP is used as a fertilizer. It can be used as fire retardant. It is also used as a yeast nutrient in wine making and brewing beer.

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UNIT 6 POLYMERS

UNIT OVERVIEW Total Periods Allotted: 14 periods

This unit mainly deals with polymers. Polymers are large molecules that are made from small units called monomers. They are generally prepared by either addition or condensation polymerization processes.

Based on their source, polymers can be divided into natural and synthetic. Natural polymers like carbohydrates, proteins and nucleic acids are essential for the existence of living organisms. Synthetic polymers are man–made and have wide applications. Even though synthetic polymers have many applications in the modern world, they have negative impacts to the environment.

This unit is divided into four subunits. The first subunit introduces polymers. It tries to explain the meaning of polymer. It also discusses classification of polymers based on source, type of polymerization and kind of monomer units.

The second subunit focuses on the two major types of polymerization processes that are applied for the preparation of polymers.

The third subunit briefly discusses the preparations and uses of synthetic polymers such as polyethylene, polypropylene, PVC, Teflon, PMMA, polystyrene, Nylon and Dacron. It also explains thermoplastics and thermosetting plastics, and environmental impacts of synthetic polymers.

The fourth subunit discusses natural polymers that occur naturally and are synthesized by living cells. The most important natural polymers that are discussed in this subunit are natural rubber, carbohydrates, and proteins.

To teach this unit, use active learning methods such as gapped lecture, group discussion, demonstration, role play and other appropriate active learning methods. Students should also be encouraged to actively participate in the teaching-learning process.

Unit Outcomes

After completing this unit, students will be able to:

- *understand the process of addition and condensation polymerization;*

- *list a variety of synthetic polymers and natural polymers and explain their differences;*
- *classify synthetic polymers as addition or condensation polymers and identify their monomers;*
- *describe monomers, properties and uses of plastics, rubbers, carbohydrates and proteins;*
- *analyze the risks and benefits of the development of and application of synthetic polymers (example, plastics) and suggest possible methods of control;*
- *describe scientific enquiry skills encountered during this unit: classifying, comparing and contrasting, communicating, asking questions, applying concepts and making generalization.*

Main Contents

6.1 Introduction to Polymers

6.2 Polymerization

6.3 Synthetic Polymers

6.4 Natural Polymers

6.1 INTRODUCTION TO POLYMERS

Periods Allotted: 1 period

Competencies

At the end of this section, students will be able to:

- *define the terms monomer and polymer;*
- *classify polymers into synthetic and natural polymers;*
- *give examples of synthetic and natural polymers.*

Forward Planning

Before you begin this lesson, it is advisable to read the contents of the topic from the Student Text Book and other reference books available in the school library. Moreover, plan a convenient way of grouping students for discussion.

Teaching Aids

Prepare a chart that shows the different classifications of polymers.

Subject Matter Presentation

For this subunit, it is possible to use brainstorming, group discussion, and presentation as active learning methods.

You can begin this subtopic by letting students brainstorm with the following terms:

- ✓ monomer
- ✓ polymer
- ✓ homopolymer
- ✓ copolymer
- ✓ natural polymer
- ✓ synthetic polymer

Then help them to form groups and discuss the note given in the Student Text Book and Activity 6.1 as well. Facilitate group discussions and clarify vague concepts for them (if any). This activity helps them to visualize how a polymer is formed. It also helps them to think at the molecular level. In this analogy, 2000 students represent 2000 ethene molecules. Since the molecular mass of the ethene (C_2H_4) is 28 amu, the molecular mass of the polyethylene is obtained by multiplying:

28×2000 amu = 56,000 amu

We know that

1 amu = 1.66×10^{-24}g

56, 000 amu = x

$$\Rightarrow x = 1.66 \times 10^{-24}\text{g} \times \frac{56,000\cancel{amu}}{1\cancel{amu}}$$

Therefore, the mass of the polyethylene molecule is 9.2960×10^{-20}g.

Finally, give them Exercise 6.1 as home work.

Assessment

Assess each student’s work throughout the subunit. This can be achieved by recording the performance of every student as usual. You can make records based on students’ performance in:

- ✓ discussing Activity 6.1
- ✓ presenting their views after discussion
- ✓ answering the questions in Exercise 6.1

By observing their performances from the record, provide them feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

1. Why are most synthetic polymers considered to be pollutants while natural polymers are not?
2. Which of the following is a copolymer?

   a. Polystyrene b. Teflon c. Natural rubber d. Nylon 66
3. Which of the following are homopolymers?

   a. Dacron b. Bakelite c. Natural rubber d. Perspex

*4. What term is used for the repeating units of a polymer?

*5. What term identifies polymers composed of identical monomer units?

Answers to Additional Questions

1. Because most synthetic polymers are non-biodegradable but natural polymers are biodegradable.
2. d
3. c & d
4. Monomer
5. Homopolymers

Answers for Exercises

Exercise 6.1

1. a. A monomer is a small unit that helps to build a polymer.

   b. A polymer is a large molecule made from monomers.

   c. A homopolymer is a polymer made from identical monomers.

   d. A copolymer is a polymer made from different type of monomers.

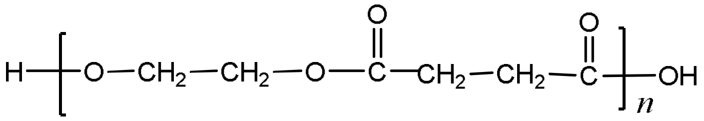
   e. A natural polymer exists naturally.

   f. A synthetic polymer is synthesized by humans.

2. You can give many examples for each type
   - **Natural polymers** – proteins, carbohydrates, nucleic acids, natural rubber etc.
   - **Synthetic polymers** – polyethylene, polypropylene, Teflon, polyvinyl chloride, nylons etc.

3.	a.	Natural	b.	Synthetic	c.	Natural
	d.	Synthetic	e.	Natural	f.	Synthetic
	g.	Synthetic	h.	Natural	i.	Natural

4.



$$H\!-\!\left[O-CH_2-CH_2-O-\overset{\overset{\displaystyle O}{\|}}{C}-CH_2-CH_2-\overset{\overset{\displaystyle O}{\|}}{C}\right]_n\!-OH$$

6.2 POLYMERIZATION

Periods Allotted: 1 period

Competencies

At the end of this section, students will be able to:

- *explain polymerization;*
- *mention the two types of polymerization;*
- *explain how addition polymerization takes place;*
- *explain how condensation polymerization takes place.*

Forward Planning

Plan how students will perform Activity 6.2 and present their answers to the class. In addition, read the relevant information in the Student Text book and refer to the answers of Exercise 6.2 in the Teachers Guide.

Subject Matter Presentation

For this subunit, it is possible to use question and answer, brainstorming, group discussion, and presentation as active learning methods.

Before you begin this lesson, review the main points of the previous lesson. Tell them to do Activity 6.2 and present their answers to the class. This activity reminds them of the functional groups that are involved in addition and condensation polymerization. Harmonize their discussions and direct them to come up with the following answers:

1. Addition reaction is the characteristic of unsaturated hydrocarbons such as alkenes and alkynes. In addition reactions, pi bonds are converted to sigma bonds. Condensation reaction occurs between compounds containing:

   acid and alcohol functional groups or acid and amine functional groups. In condensation reaction small molecules such as water are removed
2. For molecules to undergo addition polymerization, they should have double or triple bonds in their structure. Molecules that can undergo condensation polymerization should have two or more functional groups in each molecule. For example, a molecule should have two acid functional groups, one acid and one alcohol functional groups, two alcohol functional groups and etc.

Then, raise the following guiding questions to the students and tell them to answer by reading the contents of their text book.

- ✓ *What is polymerization?*
- ✓ *Can you define addition and condensation polymerization?*
- ✓ *What is the difference between addition and condensation polymerization?*

Provide them about 15 minutes to answer the above questions and encourage them to present their answers to the class. If they are not clear on the concepts, give them a brief explanation about polymerization, including addition and condensation polymerizations. Use the examples indicated in the student text book to clarify vague ideas.

Then give them Exercise 6.2 as class work and let them present their answers to the class. Finally provide feedback to them.

Assessment

Assess each student's work throughout the subunit. This can be achieved by recording the performance of every student as usual. You can make records based on students' performance in:

- ✓ discussing Activity 6.2
- ✓ presenting their views after discussion
- ✓ answering the questions in Exercise 6.2.

By observing their performances from the record, provide them with feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

*1. Which types of polymers are non-biodegradable? Why?

2. What is a characteristic of all condensation polymers?

3. What monomers are most commonly used in addition polymerization?

   a. Cycloalkanes b. Esters c. Carboxylic acids d. Alkenes

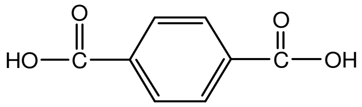
*4. What term is used to indicate a polymerization that occurs without the growth of a chain monomer by monomer, but instead occurs as monomers form dimers, dimers form tetramers, etc.?

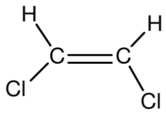
Answers to Additional Questions

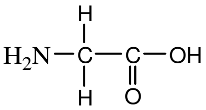
1. Addition polymers. Because the chemical bonds are very strong and cannot easily be broken by microorganisms.
2. A small molecule is lost as bonds form between the polymerizing molecules
3. d
4. step-growth or condensation polymerization

Answers to Exercises

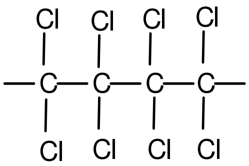
Exercise 6.2

1. a)  and $HO-CH_2CH_2-OH$

   b) 

   c) 

   $H_2N-CH_2-C(=O)-OH$

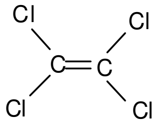
2. a. 

   $-(CCl_2-CCl_2-CCl_2-CCl_2)-$

   b. Molar mass of polymer = Number of tetrachloroethylene × molar mass of tetrachloroethylene

   $$= 3.2 \times 10^3 \times 166 \text{ g/mol} = 5.312 \times 10^5 \text{ g/mol}$$

c. The percentage of chlorine and carbon in the monomer is the same as in the polymer:



$$\% \text{ Cl} = \frac{142}{166} \times 100\% = 85.5\%$$

$$\% \text{ C} = \frac{24}{166} \times 100\% = 14.5\%$$

6.3 SYNTHETIC POLYMERS

Periods Allotted: 5 periods

Competencies

At the end of this section, students will be able to:

- *give examples of addition polymers;*
- *tell the monomers of each of the given addition polymers;*
- *explain common uses of addition polymers;*
- *give examples of condensation polymers;*
- *tell the monomers of each of the given condensation polymers;*
- *explain the uses of condensation polymers: nylon, polyester and Bakelite;*
- *describe thermoplastics and thermosetting polymers;*
- *give examples of thermoplastics and thermosetting polymers.*

Forward Planning

This subunit encompasses the synthesis and uses of addition and condensation polymers. Prior reading about the contents from the Student Text book and the reference books available in the school library is very useful. Make sure that the chemicals and apparatus for Experiment 6.1 are available. Moreover do the experiment first by yourself and check whether the experiment gives the expected results or not.

Teaching Aids

Beakers, dropping funnel, glass rod, and forceps or copper wire hooks, 0.5 M basic solution hexamethylenediamine, 0.5 M NaOH solution and adipoyl chloride.

Subject Matter Presentation

Addition polymers

For this topic, it is possible to use group discussion and presentation as active learning methods.

Students may know the effects of synthetic polymers and they should be given a chance to explore the problems of synthetic polymers and propose viable solutions. To achieve this purpose, help them to form group and do **Activity 6.3**. After discussion, facilitate their presentations to the class. Harmonize their discussions and direct them to come up with the following answers:

1. They pollute the environment and cause health risks.
2. Recycling plastics. Preparing plastics that degrade in very short periods of time and that produce nontoxic chemicals during their degradation and etc.

Then encourage them to give a brief explanation of the preparations and uses of polyethylene by reading the information in their text book.

In the subtopic of polypropylene, help them to form groups and do **Activity 6.4**. After discussion, encourage them to share their ideas with the class.

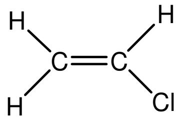
This activity lets students know the possible sources of ethene in our country. The sugar industries in Ethiopia produce molasses as a by-product, which is used for producing ethanol. Ethanol can be converted to ethene by dehydration process.

$$CH_3CH_2OH \xrightarrow{\text{Conc. } H_2SO_4} CH_2{=}\,CH_2 + H_2O$$

Students should be aware that ethanol can serve as fuel by mixing it with petrol.

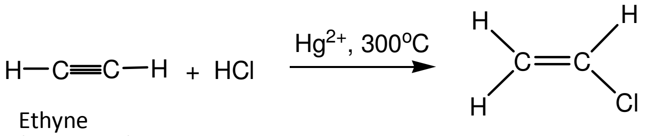
Mixing ethanol with petrol (about 5% to 95% respectively) is very common in Ethiopia and other developing countries. After summarizing preparations and uses of polyethylene and polypropylene, tell the students to form groups and address the questions in Activity 6.5. The answers are given below:

1.



Vinyl chloride

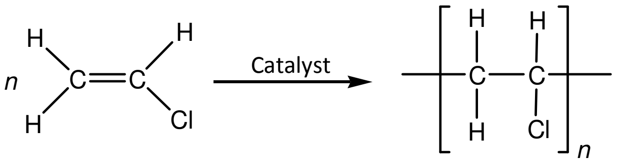
2.



Ethyne (Acetylene)

Vinyl chloride (Chloroethene)

3.



Vinyl chloride

Polyvinyl chloride (PVC)

Then give them Exercise 6.3 as homework.

Revise the main points of the previous subtopics by the question and answer method, and start these subtopics by letting students discuss in groups about the preparations and uses of polymethyl methacrylate, Teflon and polystyrene.

Encourage them to present the preparation and uses of the three polymers by assigning group representatives. Then praise the presenters and give them feedbacks on their presentations.

Encourage the students to list down materials made from polystyrene and compare the physical and chemical properties of polystyrene and PVC in Activity 6.6.

The answers for Activity 6.6 depend on the materials that students listed. Generally, materials made from Teflon have low coefficient of friction. As result, they are used to coat utensils. Materials made from polystyrene are usually used for packing other materials and are impact resistant.

Condensation Polymers

For this topic, it is recommended to use gapped lecture, group discussion, presentation, and laboratory experiment as active learning methods.

You can start this section by reiterating the key points of addition polymers that they learned in the previous section. You should emphasise the difference in the preparation of addition and condensation polymers. Provide a brief discussion about the synthesis and uses of nylon 66 and Dacron. Help them to do Experiment 6.1and prepare nylon 66 from adipoyl chloride and hexamethylene diamine. The detailed procedure is indicated in the student's text book.

Answers to the Questions in Experiment 6.1

1. Cyclohexane is a nonpolar solvent while water is a polar solvent.
2. Adipoyl chloride is more reactive than adipic acid.

Conclude this session by giving them **Exercise 6.5** as homework.

The third section discusses the preparation and uses of Bakelite, thermoplastics and thermosetting polymers. You can begin this section by giving a brief lecture (gapped lecture) about the preparation and properties of Bakelite, and the characteristics of thermoplastics and thermosetting polymers. Then let them give their answers for **Exercise 6.6** and encourage students to present their answers to the class. Praise the presenters and comment on their answers.

The last section of this subtopic explains the environmental impacts of synthetic polymers. Even though synthetic polymers have many benefits to society, they have disadvantages because they pollute the environment. In **Activity 6.7** students are assigned to collect plastics that are discarded in the environment and discuss their advantages and disadvantages.

The activity is aimed at acquainting students with practical problems of plastics and lets them propose viable solutions. Ask them to give the possible solutions that could be used to alleviate pollution of the environment by synthetic polymers. For this question, students are expected to propose recycling plastics, preparing plastics that degrade in very short periods of time and that produce nontoxic chemicals during their degradation as mitigation measures to alleviate plastic pollutions.

Then tell them to read the relevant information from their text book and answer the questions in **Exercise 6.7.** Finally conclude the subunit by recapping the main points.

Assessment

Assess each student's work continuously throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' participation in:

- ✓ discussing the activities from Activity 6.3 to Activity 6.7.
- ✓ presenting their views after discussion.
- ✓ answering oral questions raised during group discussion.

In addition, give them the exercises from Exercise 6.3 to Exercise 6.7 as class work or home work and record their achievements.

By observing their performances from the record, provide them with feedback to improve students' learning. You can also assess their performance by letting them reflect the report for Experiment 6.1.

Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit.

Additional Questions

1. Why do different polymers have different properties?
2. What are plasticisers?
3. Describe the main purposes of adding plasticisers in the synthesis of polymers.
4. Why are olefins (alkenes) good monomers for addition polymerization reactions? Examine samples of LDPE (squeeze bottles) and HDPE (milk jugs).
5. What are some of the differences in the physical properties of these substances? What is the role of a catalyst?

*6. Which of the following is correct?

a. A monomer is made from many polymers

b. A polymer is made from many monomers

c. A monomer will always dissolve in water

d. A polymer will always react with water.

*7. Most synthetic polymers are not biodegradable. This means that

a. They are expensive

b. They cannot be grown in soil

c. They will not decompose in soil

d. They only exist for a short time

*8. Polyvinyl chloride is used:

a. As a foam for fire extinguishers

b. To make leather-like materials

c. To conduct electricity in power stations

d. As reflectors in streetlights

*9. Explain the terms 'thermosetting' and 'thermoplastic polymers', with reference to molecular structure.

Answers to Additional Questions

1. They have different chemical compositions (different monomer units), different structures, different ways of being fabricated, etc.
2. A plasticiser is a substance which, when added to a material, usually a plastic, makes it flexible, resilient and easier to handle. Early examples of plasticisers include water to soften clay. Today, modern plasticisers are manmade organic chemicals, the majority of which are esters, such as adipates and phthalates
3. The electrons in the weak pi-bonds can be used to form strong sigma bonds to other monomer units.
4. LDPE – more transparent, flexible, waxy. HDPE – more opaque, rigid, non-waxy
5. A catalyst reduces the energy of activation for a reaction by providing an alternative pathway. In this way, it speeds up the reaction and allows it to proceed under milder conditions

6. b 7. c 8. b

9. Thermosetting: once 'set' these plastics cannot be reheated to soften, shape and mold. The molecules of these plastics are cross linked in three dimensions and this is why they cannot be reshaped or recycled. The bond between the molecules is very strong.

Thermoplastics: these plastics can be re-heated and therefore re-shaped in various ways due to long chain polymers that are not inter-connected. They become moldable after reheating as they do not undergo significant chemical change. Reheating and shaping can be repeated. The bond between the molecules is weak and becomes weaker when reheated, allowing reshaping. These types of plastics can be recycled.

Answers to Exercises

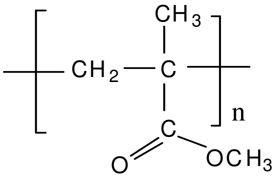
Exercise 6.3

1. a. C_2H_4, $H_2C=CH_2$ b. C_3H_6, $CH_2=CH\ CH_3$
   c. C_2H_3Cl, $CH_2=CHCl$
2. a. Polyethylene is used for making squeeze bottles, plastic wrapping, and electrical insulation.
   b. Polypropylene is used for making food containers, ropes and carpets.
   c. Polyvinyl chloride is used for making pipes, leather – like coats, wallpaper and floor tile.

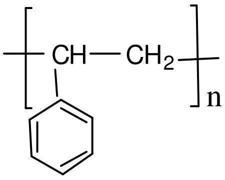
3. Pipes, raincoats, floor tile and etc.
4. HDPE is a linear polymer and LDPE is a branched polymer.

Exercise 6.4

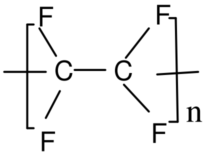
1. a. Polymethyl metharylate,



b. Polystyrene,



c. Polytetrafluoroethylene,



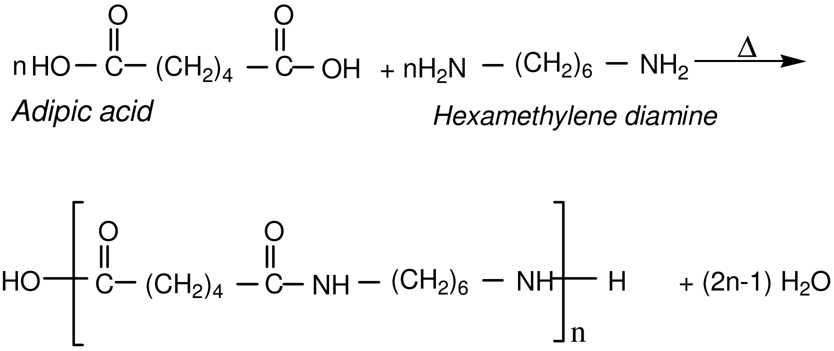
2. a. Teflon is mainly used to coat cooking utensils.

b. Polystrene is used for making packaging materials, CD and DVD cases, and architectural models.

c. Plexiglass or polymethyl methacrylate is used to make airplane windows and as a glass substitute.

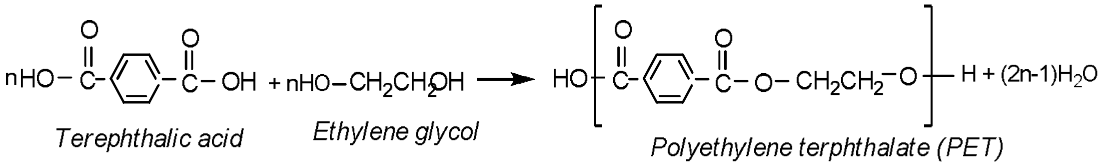
Exercise 6.5

1. a)

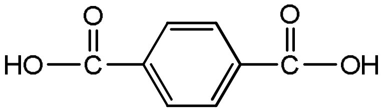
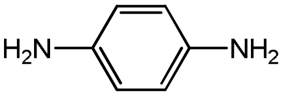


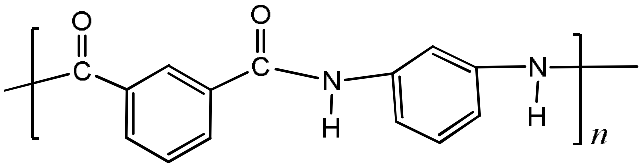
$$nHO-\overset{O}{\overset{\|}{C}}-(CH_2)_4-\overset{O}{\overset{\|}{C}}-OH + nH_2N-(CH_2)_6-NH_2 \xrightarrow{\Delta}$$

$$HO-\left[\overset{O}{\overset{\|}{C}}-(CH_2)_4-\overset{O}{\overset{\|}{C}}-NH-(CH_2)_6-NH\right]_n-H \quad + (2n-1)\,H_2O$$

b) 

2. Nylons are used to make parachutes, ropes and stockings. Dacron is mainly used to make synthetic fibres.

3.  and 

4. 

5. Both the amine functional groups and acid groups in Kevlar have 1, 4- or para positions, but in Nomex their positions are 1, 3- or meta positions.

Exercise 6.6

1. The monomers of Bakelite are formaldehyde and phenol.
2. Bakelite is used for making rotary–dial telephones, billiard balls, chess, pieces and porcelain.
3. Thermoplastics soften on heating and can be moulded into different shapes. For example polyethylene is a thermoplastic and can be recycled. Thermosetting plastics cannot be softened once moulded. Bakelite is a thermosetting plastic and cannot be recycled.
4. In thermoplastics, the bond between the molecules is weak and become weaker when reheated, allowing reshaping. These types of plastics can be recycled.

In thermosetting plastics, the molecules are cross linked and the bond between the molecules is very strong. This is why they cannot be reshaped or recycled.

Exercise 6.7

1. Major problems caused by plastics are environmental pollution and health risks
2. Yes, the higher the durability, the higher the extent of pollution. This is because durable plastics for a last long time, polluting the environment.

6.4 NATURAL POLYMERS

Periods Allotted: 7 periods

Competencies

At the end of this section, students will be able to:

- *describe natural rubber;*
- *tell the monomer of natural rubber;*
- *explain vulcanization;*
- *explain the use of natural rubber;*
- *give three examples of synthetic rubber;*
- *explain the uses of synthetic rubbers;*
- *compare and contrast synthetic and natural rubber;*
- *define the term carbohydrate;*
- *describe monosaccharide;*
- *draw the structures of monosaccharides;*
- *give examples of monosaccharides;*
- *describe disaccharides;*
- *draw the structures of disaccharides;*
- *describe polysaccharides;*
- *draw the structure of starch and cellulose;*
- *explain the differences between starch, glycogen and cellulose;*
- *define the terms amino acid, peptide and protein;*
- *describe the structure of amino acids;*
- *explain how proteins are formed;*
- *list types of protein.*

Forward Planning

Read the relevant information in the Student Text and reference books available in the school library. In addition, investigate the answers of the exercises and activities before you start the subtopic.

Subject Matter Presentation

For this topic, it is advisable to use gapped lecture, group discussion, and presentation as active learning methods.

Begin this subunit by allowing students to perform **Activity 6.8.** Then, let them discuss in groups, and share their ideas with the rest of the class. This activity is aimed at inspiring students with the presence of natural polymers that are found in our body and in our surroundings. Hence, students are expected to list down polymers found in the environment and in the human body and classify them as condensation or addition polymers. Most natural polymers such as starch, cellulose, glycogen, proteins and nucleic acids are condensation polymers. However, natural rubber and terepenes are addition polymers.

The first subtopic discusses natural and synthetic rubber. This subtopic may be new for students. Start it by giving a brief explanation about natural and synthetic rubber. Emphasis should be given to:

- the monomers and polymers of natural and synthetic rubber;
- the difference in properties of natural and synthetic rubber;
- the application of vulcanization.

All these concepts are discussed in the student's text book.

Then give them Exercise 6.8 as class work and encourage them to present their answers to the class. Provide feedback to the students.

Carbohydrates

For this subtopic, you can use brainstorming, independent work, and gapped lecture as active learning methods.

Students may be familiar with carbohydrates from work in lower grades or in biology lessons. Give them the following questions for brainstorming:

- How are carbohydrates produced by plants?
- What are the elements found in carbohydrates?
- Define carbohydrates and monosaccharide.
- What are aldoses and ketoses?

Then give them 10 minutes to answer the above questions by reading the contents of carbohydrates and monosaccharides in their text book.

If they are not clear on the concepts, provide brief explanations about:

- ✓ the definition of carbohydrates, monosaccharide, aldoses and ketoses;
- ✓ the open chain structures of glucose, galactose, fructose, and ribose;
- ✓ the structural similarity and differences of glucose, and fructose, ribose and deoxyribose;
- ✓ the cyclic structures of glucose and fructose.

All these concepts are found in the student text.

End this section by giving Exercise 6.9 as homework.

Disaccharides

For this subtopic, it is advisable to use brainstorming and gapped lecture as active learning methods.

Students may be familiar with disaccharides from their biology lessons. So, you can start this subtopic by letting students brainstorm about disaccharides.

- ✓ List down disaccharides and their monomers.
- ✓ How do disaccharides form?

Then, provide a brief lecture (gapped lecture) about the structure and uses of maltose, cellobiose, lactose and sucrose. Give them **Exercise 6.10** as class work and encourage them to present their answers to the class. Then conclude this subtopic by giving feedback to the students.

Polysaccharides

For this subtopic, it is possible to use brainstorming, group discussion, presentation, and gapped lecture as active learning methods

You can start this subtopic by raising the following questions to the class:

- ✓ What does polysaccharide mean?
- ✓ What are the monomers of cellulose, starch and glycogen?
- ✓ Is there any structural difference between cellulose, starch and glycogen? What is that difference?

Let them read their text book, discuss in groups and answer the questions. Then use gapped lecture to explain the structure and uses of cellulose, starch and glycogen.

Give them Activity 6.9. This activity reminds them that human beings have learned a lot from chemistry of natural polymers and they have applied what they have learned to produce synthetic polymers. The first question helps them to associate the synthesis of polysaccharides by plants such as cotton with the synthesis of synthetic fibres by humans using condensation reaction. Synthetic fibres such as polyesters, nylons and etc can be taken as examples.

The answer for the second question is the following: Cotton burns quickly to give the odour of burnt paper and leaves grey ash after combustion. Wool and natural silk burn slowly to give the odour of burnt feather. A black ball is formed after combustion which can be crushed to powder. People usually use this technique to differentiate wool from synthetic fibres.

You may give them Exercise 6.11 as class work and encourage them to present their answers to the class. Conclude this subtopic by giving feedback to students.

The next subtopic focuses on amino acids.

Proteins

For this subtopic, it is possible to use gapped lecture and brainstorming as active learning methods.

You can start this section by giving a brief explanation of the meaning of protein and amino acids. In your discussion, you should emphasize the structure of α–amino acids and the classification of amino acids based on side chain (non-polar, polar neutral, polar acidic and polar basic amino acids). Then give them **Exercise 6.12** as class work and let them present their answers to the class. End this session by giving feedback to students.

The last subtopic explains the physical properties of amino acids and peptide bond formation. You may start this section by reviewing the main concepts of the previous lesson. Then explain:

- why amino acids have high melting points;
- zwitterions and the charge of amino acids at low pH, pH = 7, and high pH;
- peptide or amide bond formation;
- the difference between polypeptide and proteins.
- the different types of proteins and their biological functions.

Then help them to form groups and discuss the questions in **Activity 6.10**. This activity acquaints them with the chemistry of amino acids and peptide bond formation. Facilitate group discussions by clarifying vague concepts for them. After discussion, encourage them to share their ideas to the rest of the class.

Harmonize their discussions and help them to reach the following answers:

1. Since they contain both acid and base functional groups in the same molecule, the acid donates proton and the base accepts the proton.
2. The number of tripeptides is 3! = 3 x 2 x 1 = 6
3. The amine functional group of one amino acid molecule reacts with the acid functional group of the other amino acid to from amide or peptide bond. This continues until proteins are formed.

4. Condensation polymerization.
5. All proteins are made from the 20 amino acids. However, all proteins do not have the same sequence of amino acids. As a result, different proteins have different functions.

Conclude this subtopic and Unit as well by summarizing the main points. Moreover, give them Exercise 6.13 as homework and ask the students to submit the answers of the review questions of the Unit.

Assessment

Assess each student's work throughout the subunit. This can be achieved by recording the performance of every student. You can make records based on students' performance in:

- ✓ discussing the issues from Activity 6.8 to Activity 6.10
- ✓ presenting their views after discussion
- ✓ answering the questions from Exercise 6.8 to Exercise 6.13.

By observing their performances from the record, provide them with feedback to improve students' learning. In addition, you can use self assessment and peer assessment methods to enhance students' performance. Appreciate students working above the minimum required level and encourage them to continue working hard. For low achievers, identify their learning difficulties and help them to achieve the minimum required level for this subunit. You can also use the review questions of this unit as part of summative assessment.

Additional Questions

1. You probably observed that some people do not like milk. If they consume dairy products, they will experience diarrhea. This problem is called lactose intolerance. Explain why this problem occurs.
2. Protein shape, function and amino acid sequences are interrelated. Which determines which?
3. On heating with concentrated HNO_3, proteins give yellow colour. This test is called Xanthoprotic test. What does this test indicate?

*4. Name the different kinds of elements found in carbohydrates.

*5. What is the type of linkage that chemically connects the subunits of carbohydrates?

*6. What functional groups are found in amino acids?

*7. What element is used to vulcanize rubber?

*8. Which type of polymer is formed from each of the following monomers?
a. Isoprene b. Chloroprene c. Glucose d. Amino acids

Answers to Additional Questions

1. Lactase is an enzyme that specifically breaks the β - 1, 4 – glycosidic linkage of lactose. Some people lose their intestinal lactase when they become adults; they are then no longer able to digest lactose. Consequently, when they drink milk, the undegraded lactose causes digestive problems such as bloating, abdominal pain, and diarrhea. These problems occur because only monosaccharides can pass into the bloodstream, so lactose has to pass undigested into the large intestine
2. The protein's amino acid sequence determines its structure, which in turn determines its function.
3. This test gives yellow colour if proteins contain aromatic amino acids such as phenylalanine, tyrosine and tryptophan.
4. C, H & O
5. Glycosidic linkage
6. Carboxyl and amine
7. sulphur
8. a. Natural rubber c. Polysaccharide
   b. Synthetic rubber. d. Protein

Answers to Exercises

Exercise 6.8

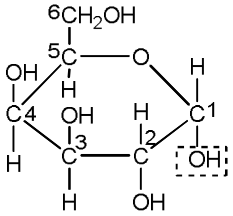
1. a. Isoprene or 2 – methyl -1, 3- butadiene.
   b,. Chloroprene or 2 – chloro -1, 3 – butadiene.
   c. Styrene and 1, 3 – butadiene.
   d. 2-Methylpropene and isoprene.
2. Vulcanization is hardening of rubber by heating rubber with sulphur. Vulcanized rubber has high tensile strength and is hard, non-sticky and insoluble in solvents.
3. Neoprene is used for making hoses for chemicals, petrol and oil. Styrene-butadiene rubber (SBR) is used for the manufacture of tires.

Exercise 6.9

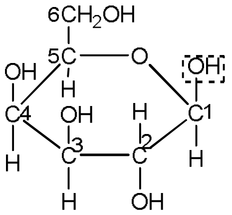
1. Carbohydrates are polyhdyroxy ketones, polyhdyroxy aldehydes or compounds that yield such substances upon hydrolysis.

2. a. Aldotetrose b. Ketopentose
   c. Aldotriose d) Ketotriose

3. Galactose

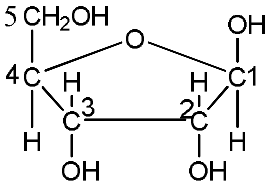


α - *Galactopyranose*

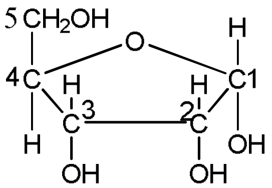


β *-Galactopyranose*

Ribose



β- *Ribofuranose*



α- *Ribofuranose*

4. Ribose ($C_5H_{10}O_5$), glucose ($C_6H_{12}O_6$), galactose ($C_6H_{12}O_6$), and fructose ($C_6H_{12}O_6$)
5. It is a sugar that is oxidized by Benedict or Fehling solution. All monosaccharides are reducing sugars.

Exercise 6.10

1. a. Glucose and fructose

   b. Glucose and glucose

   c. Galactose and glucose

   d. Glucose and glucose

2. a. α, β - 1, 2 – glycosidic linkage

   b. α - 1, 4 – glycosidic linkage

   c. β - 1, 4 – glycosidic linkage

   d. β - 1, 4 – glycosidic linkage

3. All have a molecular formula of $C_{12}H_{22}O_{11}$.

4. a. A lemon juice contains citric acid. Hence it is used as a catalyst in the hydrolysis of sucrose.

b. The hydrolysis of sucrose results in the formation of glucose and fructose. Fructose is the sweetest of all sugars. As a result, the product mixture is sweeter than the starting sucrose solution.

Exercise 6.11

1. a. Both glycogen and amylopectin have α - 1, 4 – and α - 1, 6 – glycosidic bonds but glycogen is more branched than amylopectin.

   b. Both amylose and amylopectin are found in starch, but amylose is a linear polymer and amylopectin is a branched polymer.

   c. Amylose and cellulose are the polymers of glucose, but the glycosidic linkages in amylose are α - 1, 4. Those in cellulose are β - 1, 4.

   d. Both amylose and glycogen have α - 1, 4 – glycosidic linkages but glycogen has also α - 1, 6 – linkages. As a result amylose is a straight chain polymer of glucose while glycogen is a branched chain polymer of glucose.

2. i. A, and B ii. B, and C iii. A and B iv. B, and D

3. Humans have enzymes to digest starch, but they do not have the enzyme cellulase to break β -1, 4 – glycosidic linkages in cellulose.

Exercise 6.12

1. a. α - amino acid

   b. β - amino acid

   c. β - amino acid

   d. α - amino acid

2. a. Nonpolar amino acids

   You can give two of the following amino acids.

   Glycine, Alanine, Valine, Leucine, Proline, Isoleucine, Methionine, and Phenylalanine.

   b. Polar neutral amino acids

   Serine, Cysteine, Theronine, Asparagine, Glutamine, Tyrosine, and Tryptophan.

   c. Polar acidic amino acids

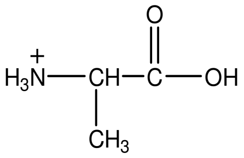
   Aspartic acid and Glutamic acid.

d. Polar basic amino acids

Histidine, Lysine, and Arginine

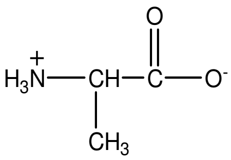
Exercise 5.13

1. i)



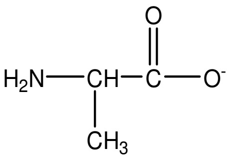
$\overset{+}{H_3N}-\underset{\substack{|\\CH_3}}{CH}-\overset{\substack{O\\\|}}{C}-OH$

ii)



$\overset{+}{H_3N}-\underset{\substack{|\\CH_3}}{CH}-\overset{\substack{O\\\|}}{C}-O^-$

iii)

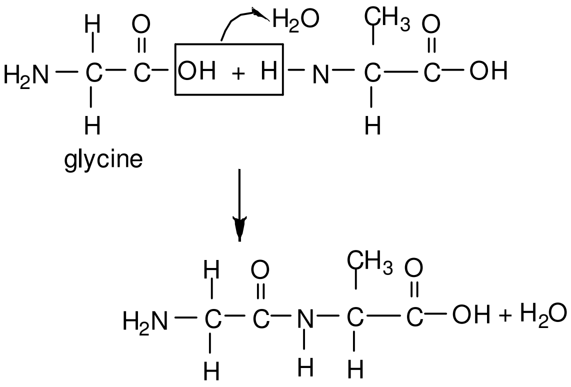


$H_2N-\underset{\substack{|\\CH_3}}{CH}-\overset{\substack{O\\\|}}{C}-O^-$

2. Carboxylic acid and amine functional groups.

3. Amino acids are building blocks of both peptides and proteins. Generally peptides contain fewer amino acid residues compared to proteins.

4.



Answers to Review Questions

Part I

1.	C	2.	D	3.	B	4.	D	5.	A
6.	B	7.	D	8.	C	9.	D	10.	B
11.	D	12.	B	13.	A	14.	C	15.	C
16.	D	17.	C	18.	A	19.	B	20.	C
21.	C	22.	C	23.	C	24.	D		

Part II

25. a. A monomer is a small unit that helps to form a polymer.

b. A polymer is a large unit formed by joining of monomer units.

c. Carbohydrates are polyhydroxy ketones, polyhydroxy aldehydes, or compounds that give these substances upon hydrolysis.

d. Amino acid contains amino and carboxylic acid functional groups.

e. Polypeptide generally contains amino acid residues from two to fifty.

f. Protein is a polypeptide that contains amino acid residues greater than 50.

26. a. Synthetic polymer
b. Synthetic polymer
c. Natural polymer
d. Synthetic polymer
e. Natural polymer
f. Synthetic polymer
g. Synthetic polymer
h. Synthetic polymer
i. Natural Polymer
j. Synthetic polymer

27. a. Addition polymer
b. Addition polymer
c. Condensation polymer
d. Addition polymer
e. Condensation
f. Addition
g. Addition
h. condensation
i. Condensation
j. Condensation

28. In addition polymerization multiple bonds (double or triple bonds) are involved, but in condensation polymerization, carboxylic acids, amines or alcohols are commonly involved.

29. a. Polypropylene is used for making ropes, carpets and food containers.

b. Polyvinyl chloride is used for making pipes, leather–like materials, shoes wallpaper and floor tile.

c. The major use of Teflon is for coating utensils.

d. Polystyrene is used for making packing materials, CD and DVD cases and architectural models.

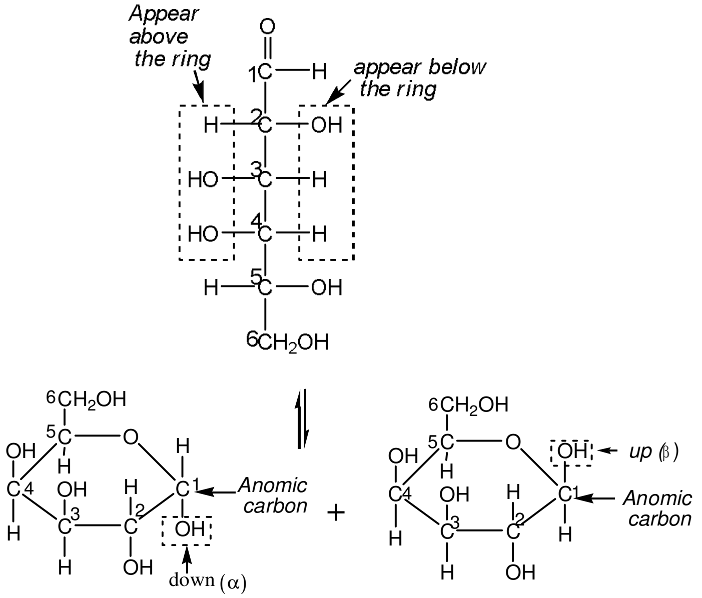
e. Bakelite is used for making rotary – dial telephones, billiard balls and porcelain.

f. Nylons are mainly used for making parachutes, clothes, stockings and ropes.

g. Plexiglas or PMMA is mainly used to make airplane windows and as a substitute for glass.

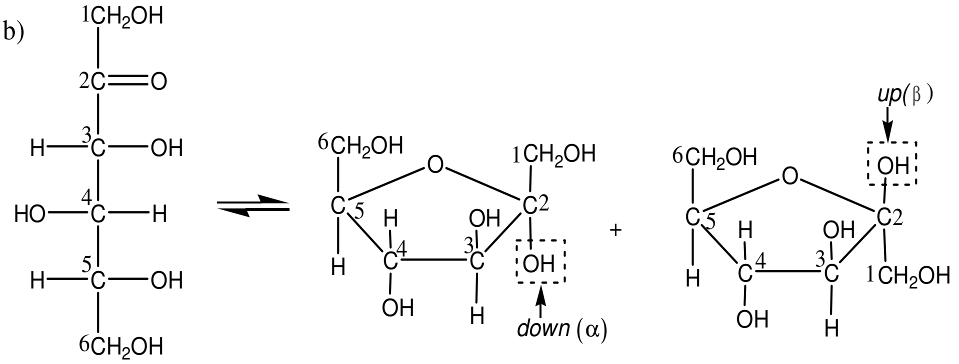
30 a)

Galactose



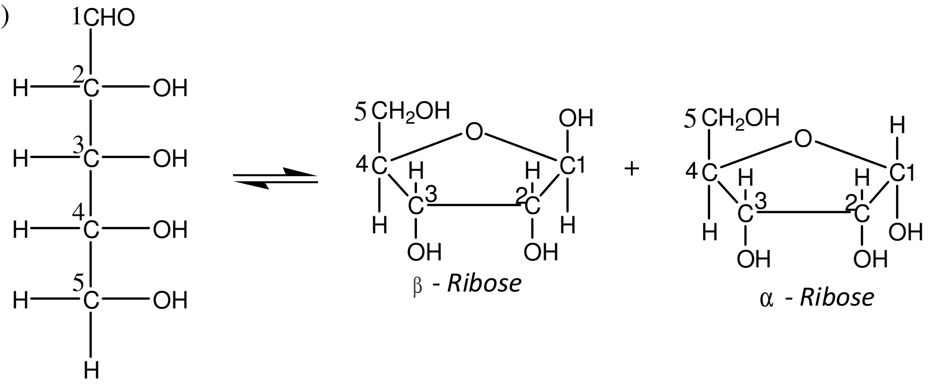
α - *Galactopyranose* β -*Galactopyranose*

b)



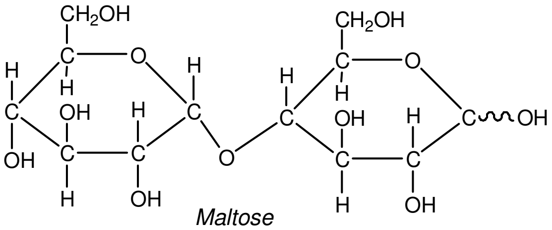
Fructose α - *Fructofuranose* β - *Fructofuranose*

c)



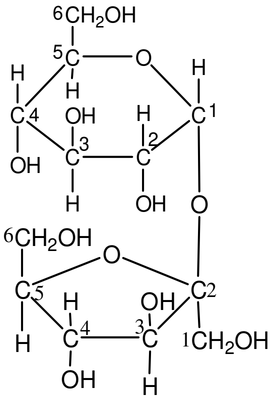
Ribose

d)



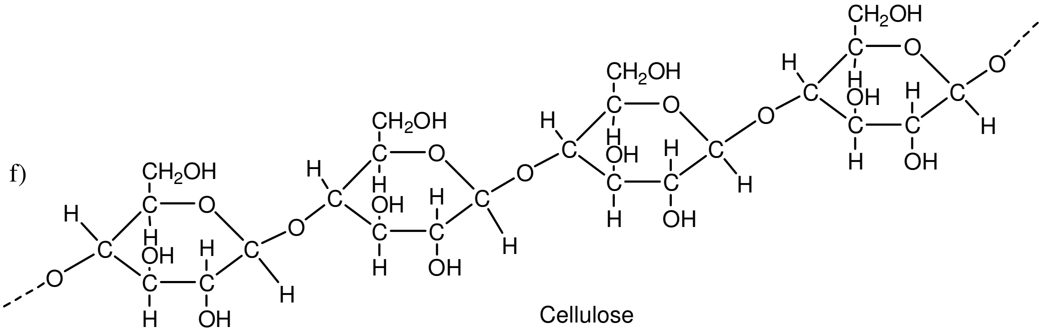
Maltose

e)



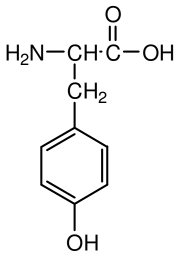
Sucrose

f)



Cellulose

g)



Phenylalanine

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MINIMUM LEARNING COMPETENCIES (MLCs)

No	Area of Competency	MINIMUM LEARNING COMPETENCIES (MLCs)
1	**CHEMICAL REACTION**	**The students should be able to:** • define the terms mixture, homogenous and heterogeneous mixtures, solute , solvent, solution • distinguish between homogeneous and heterogeneous mixtures • explain different types of solutions and give examples for each • define heat of solution , salvation energy and hydration energy • explain the solution process and how heat of solution is influenced by the interparticle interaction forces • explain formation of saturated and supersaturated solutions • define rate of dissolution and discuss the factors that affect it • define solubility and describe the factors that affect it • conduct an experiment to determine solubility of table salt and sugar • state Henry's law and use it to calculate concentration of gaseous solute in a solution • define concentration of a solution, mass percentage of a solute in a solution, mole fraction , molarity, molality, equivalent mass, number of equivalents and normality • calculate the mass percentage and mole fraction of a solute in a solution • prepare molar , normal and molal solutions of different substances • calculate molarity, normality and molality of a solution from a given information • explain dilution process and calculate the volume or concentration changes during dilution of solution • do calculations involving solutions to determine the number of moles , masses or volumes of reactants and products in ionic reactions • analyze ionic reactions and write net ionic equations • List the important properties of solvents that are affected by the formation of a solution • State Raoult's law • explain the changes in the colligative properties of solvent when a certain solute is added • describe the Vant Hoff's factor • calculate the vapor pressure, the boiling point and the freezing point of a solvent after a certain amount of solute is added • define osmosis and osmotic pressure of a solution • calculate the osmotic pressure of a solution • compare and contrast change in colligative properties of

		electrolytic and nonelectrolytic solution • define acid and base by Arrhenius , Bronsted-Lowry and Lewis concepts and give examples of each • define chemical thermodynamics, systems, state functions, path function and non-function • give examples of systems, classify them into open, closed and isolated systems and explain them • distinguish between intensive and extensive properties • explain internal energy, heat and work in relation to the concepts of thermodynamics • state and explain first law of thermodynamics • calculate change in internal energy of a system from a given information • define enthalpy change , standard state, standard molar enthalpy of combustion, standard molar enthalpy of formation and standard molar enthalpy • carry out an activity to measure standard molar enthalpy of neutralization • state Hess's law and apply it to solve problems on enthalpy changes of chemical reactions • explain bond energy and calculate it in a given chemical reaction • calculate the standard enthalpy changes of a reaction from given enthalpy changes of reactants and products • explain entropy and entropy changes • calculate the entropy change from the given standard entropies of substances • state and explain the second law of thermodynamics • calculate the entropy changes from the given enthalpy change of the system and absolute temperature • explain free changes and standard free change • calculate the standard free energy change from a given standard free energies of reactants and products • describe the relationship between standard free change, standard enthalpy change and standard entropy change of a reaction • determine the spontaneity of a reaction • define redox reactions, oxidation and reduction • describe the oxidizing and reducing agents • identify the species that are oxidized and reduced in a given redox reaction • balance a given redox reaction using oxidation number method and ion-electron method • distinguish between molten electrolytes and aqueous electrolytic solutions

		• draw labeled diagram of an electrolytic cell • define preferential discharge and explain factors that affect it • describe the reactivity of a metal from its position in the activity series • describe the effect of nature of the ions, concentration of the ions, types of the electrodes on the electrolysis of aqueous solutions of dilute sulphuric acid and sodium hydroxide , on the electrolysis of dilute and concentrated solutions of sodium chloride and on the electrolysis of dilute sulphuric acid, dilute sodium hydroxide, concentrated and dilute sodium chloride and copper sulphate solutions respectively • state Faraday’s first and second law of electrolysis and write mathematical expression of them • do calculations related to Faraday’s first and second law of electrolysis • mention industrial applications of electrochemistry • define voltaic cell, salt bridge, electrode potential and cell potential • draw, label and construct Zn-Cu voltaic cell • measure the cell potential of Zn-Cu cell using voltmeter • explain how standard electrode potential is measured and calculate cell potential • decide whether a given redox reaction is spontaneous or not • explain the effect of concentration on cell potential • mention different types of voltaic cells and give examples of each • distinguish between primary and secondary cells • identify the cathode, anode and electrolyte of a given voltaic cell • compare and contrast electrolyte and voltaic cells • explain metallic corrosion in terms of redox • reaction , its negative effects and different methods of prevention • describe the forms of occurrence of metals • define metallurgy and explain the major steps in metallurgical processes • explain the chemical properties, the manufacture and uses of sodium • explain the chemical properties , the manufacture and uses of calcium • explain the chemical properties , the manufacture and uses of tin • explain the chemical properties , the manufacture and uses of lead • explain the chemical properties , the manufacture and uses of zinc • explain the chemical properties , the manufacture and uses of chromium

2	**CARBOXYLIC ACIDS, ESTERS, FATS, OILS AND POLYMERS**	• define monomer and polymer • classify polymers into synthetic and natural polymers give examples for each • explain polymerization and mention two types of polymerizations • explain how addition and condensation polymerization take place • give examples of addition polymers and tell their monomers • give examples of condensation polymers and tell their monomers • explain the common uses of addition and condensation polymers • describe thermoplastics and thermo set polymers and give examples of each • describe natural rubber and tell its monomers • explain vulcanization • explain the use of natural rubber • give three examples of synthetic rubbers and explain their uses • compare and contrast synthetic and natural rubber • define carbohydrate • describe monosaccharide, draw its structure and give examples • describe disaccharide, draw its structure and give examples • describe polysaccharide, draw its structure and give examples • explain the difference between starch, glycogen and cellulose • define amino acid, peptide and protein • describe the structure of amino acids • explain how proteins are formed • list types of proteins

Federal Democratic Republic of Ethiopia
Ministry of Education

Chemistry Syllabus, Grades 12

2009

Table of Contents

INTRODUCTION

According to the Educational Training Policy of Ethiopia, the second cycle of the secondary education training will enable students to choose subjects or areas of training which will prepare them adequately for higher education for the world of work.

On this basis this revised chemistry syllabus is prepared in such a way that it can address those areas of study that chemistry students could attend at higher education. Students of chemistry are expected to join chemistry, physics, biology, medicine, pharmacy, earth science, engineering, agriculture the likes. It is believed that there are grains of contents that lay ground for the above mentioned fields of studies.

The revision of the curriculum was based on the new curriculum framework of Ethiopian schools assessments made in selected schools of the country.

Generally the process of the revision involved incorporating the comments given by the students the teachers during the assessment, meeting the requirements of the framework keeping international standards.

The major points that were considered in the revision were:

- Making the content overload appropriate to the period allotment;
- Making content difficulties appropriate to the grade level;
- Emphasise active learning approaches;
- Reducing unnecessary repetitions;
- Restructuring the curriculum to competency based;
- Strengthening the bridging role of the level to higher education;
- Integrating some technological, agricultural health concepts.

To achieve the above mentioned purposes, some restructuring works have been done all through grades 7- 12. Consequently, both grades 11 and 12 are made to have 6 units each.

This reduction of the contents is believed to facilitate active learning process.

To mention few typical examples that has been made on this revision:

- In both grades 11 and 12, enough practical activities are included to implement active learning approaches,
- In grade 11, the title “phase equilibrium” is introduced in unit 5(Chemical Phase Equilibrium) to deepen students’ knowledge on the properties of the three states of matter their phase changes,
- To solve the problems of content overload, the past very broad wide four units of grade 12, s-block metals, p- block metals, p-block non-metals d-block metals are distributed integrated to different grade levels reduced here to one unit under the title “The Elements in Nature Industry”. This unit treats only the occurrence abundance of elements, recycling of carbon, nitrogen phosphorus in nature, production of some metals such us sodium, calcium, tin, lead, zinc chromium, metalloid, which is silicon , compounds of some non-metals such us nitric acid, sulphuric acid Diammonium Phosphate(DAP).

The format of the syllabus is different from the traditionally used one. There are only three columns of competency, content suggested activities respectively in the syllabus below which comes the assessment row.

In the assessment the minimum learning competencies for students working at the minimum requirement level are listed to evaluate their performances. It also suggests assistance to be made for students working below above the required levels.

Assessment is done through continuous process; however, specific assessment techniques are selected in order to collect information about how well students are achieving the competencies. The assessment techniques used at any particular time depends on what facility with the knowledge, skill, or process the teacher wants the student to demonstrate. The appropriateness of the techniques therefore results on the content, the instructional strategies used, the level of development of the students what is to be assessed. The environment culture of the students must also be considered.

In the assessment, the statement **"minimum requirement level"** should not mislead should be understood as the **"standard level"**. Students working at the standard level are expected to achieve the competencies set for the grade level successfully.

Various assessment techniques are listed below. The techniques listed are meant to serve only for reference, since the teacher exercises professional judgement in determining which technique suit the particular purpose of assessment.

Correlating Instruction, Evaluation, Science Goals

Instructional Strategies	Some Important Active Learning Methods for Science	Some Corresponding Assessment Techniques
Direct	• Demonstrations	• Group/Individual (Peer/Self): Performance Assessments • Short-Answer Quizzes and Tests
Indirect	• Concept Mapping/Formation/ Attainment • Inquiry • Problem Solving	• Individual/Group: Presentations • Oral Assessments • Performance Assessments • Written Assignments
Experiential	• Conducting Experiments • Field Observations and Trips • Model Building • Simulations	• Group/Individual: Performance Assessments; • Written Assignments; • Peer/Self: Oral Assessments • Technical Skills
Independent Study	• Reports • Homework • Research Projects	• Performance Assessments • Portfolios • Presentations • Quizzes • Written Assignments
Interactive	• Brainstorming • Co-operative Learning Groups • Discussion • Laboratory Groups	• Group/Peer: Oral Assessments • Written Assignments

To reinforce utilization of the syllabus other curriculum materials are also prepared to accompany it. These materials are: flowchart, minimum learning competency, student textbook, teacher's guide, practical activities manual, students' workbook.

The syllabus for Grades 11 12 were revised by six consultants, teachers curriculum experts.

Consultants

Ato Chala Regassa (AAU)

Dr. Tetemke Mehari (AAU)

Teachers from Addis Ababa

Ato Guta Degefa (Medhanealem senior Secondary School)

Ato Mekonnen Legesse (Menilik II Senior Secondary School)

Curriculum Experts from Ministry of Education

Ato Nega Gichile

Ato Tesfaye Jinore

Allotment of Periods for Units Sub-units of Chemistry Grades 12

Unit	*Sub-unit*	*Number of Periods*	
		Sub-unit	*Total*
Unit 1: Solutions	1.1 Homogenous and heterogeneous mixtures	1	30
	1.2 Types of solutions	2	
	1.3 The solution process	5	
	1.4 Solubility as an equilibrium process	4	
	1.5 Ways of expressing concentrations of solutions	7	
	1.6 Preparation of solutions	2	
	1.7 Solution stoichiometery	3	
	1.8 Describing reactions in solution	1	
	1.9 Colligative properties of solutions	5	
Unit 2: Acid-base Equilibrium	2.1 Acid base concepts	5	26
	2.2 Ionic equlibiria of weak acids and bases	9	
	2.3 Common ion effect and buffer solutions	4	
	2.4 Hydrolysis of salts	2	
	2.5 Acid base indicators	6	
Unit 3: Introduction to Chemical Thermodynamics	3.1 Common thermodynamic terms	2	12
	3.2 First law of thermodynamics and some thermodynamic quantities	3	
	3.3 Thermo chemistry	4	
	3.4 Entropy and second law of thermodynamics	3	
Unit 4: Electrochemistry	4.1 Reduction- oxidation reacting	5	21
	4.2 Electrolysis of aqueous solutions	6	
	4.3 Quantitative aspects of electrolysis	3	
	4.4 Industrial application of electrolysis	2	
	4.5 Voltaic cells	5	
Unit 5: Some Elements in Nature and Industry	5.1 Some elements in nature	5	17
	5.2 Some elements in industry	12	
Unit 6: Polymers	6.1 Polymerization	2	14
	6.2 Synthetic polymers	5	
	6.3 Natural polymers	7	

General Objectives of Grade 12 Chemistry

To develop understanding and acquire knowledge of:

- types, properties and processes of formation of solutions;
- different concepts of acids and bases, their properties and qualitative and quantitative expressions;
- energy changes in physical and chemical processes and laws governing them;
- fundamental concepts related to the interconversion of chemical and electrical energy;
- qualitative and quantitative aspects of electrochemistry and its applications;
- occurrence, abundance and metallurgical processes of some useful elements;
- analyzing the risks and benefits of the development and application of synthetic polymers and suggesting possible methods of control.

To develop skills and abilities of:

- preparing solutions of specific concentration and solving quantitative problems involving solutions;
- solving equilibrium problems involving acids and bases;
- predicting weather a solution of a specific salt will be acidic, alkaline or neutral;
- calculating energy changes in chemical reactions;
- determining spontaneity of chemical reactions;
- calculating problems related to electrolysis;
- constructing model galvanic cell;
- designing and conducting simple experiments relevant to their level.

To develop habits and attitudes of:

- appreciating the roll of some chemical processes such as buffer systems in balancing natural phenomena in our lives;
- practicing the chemical and physical methods of preventing corrosion;
- realising that elements are cycling in nature;
- being confident and independent;
- relating a scientific knowledge to everyday applications.

Unit 1: Solutions *(30 periods)*

Unit out comes: Students should be able to:

- know the types of solutions;
- understand the solution formation process, the rate of solution , the heat of solution & solubility;
- describe the dependence of solubility on temperature & pressure of solution;
- know how to solve problems involving concentration of solutions & express the result in various units;
- describe, using the concept of equilibrium, the behavior of ionic solutes in solutions that are unsaturated, saturated and supersaturated;
- prepare solutions of required concentration by dissolving a solute or diluting a concentrated solution;
- understand the relationship between the colligative properties of solution & know how to solve problems involving colligative properties of solution;
- describe scientific enquiry skills along this unit: observing, classifying, comparing & contrasting, communicating, measuring, asking questions, drawing conclusion, applying concept and problem solving.

Competencies	*Contents*	*Suggested Activities*
Student's will be able to: • Define the terms mixture, homogenous and heterogeneous mixtures, solute, solvent, solution. • distinguish between homogenous and heterogeneous mixtures • describe suspension and colloids	**1. Solutions** **1.1 Homogeneous and heterogeneous mixtures** *(1 period)* • Some basic definitions - mixture, homogeneous and heterogeneous mixtures, solute, solvent, solution • Suspension and colloids	Students should be aware by now that understanding and using proper terminologies from the very beginning helps master basic concepts in science. Student should know that mixtures are divided in to three categories: suspensions, colloids, and solutions. A suspension is a dispersion of fine solid particles in a liquid or gas, removable by filtration. It could also separate into its components under the influence of gravity. A mixture of sand and water is an example. Students should know that in a solution, the particles of the solute are of the size of individual atoms, molecules or ions. Because of their small size they mingle with the particles of the solvent and never separate, no matter how long the solution is allowed to stand. Between these two extremes are mixtures called colloids. In a colloid, the particle size is much larger than in a solution, but not so large that they separate on standing. Ask students to dissolve sand, Gentian Violet and bile in water in three different flasks. Set them aside and see if they become any clearer on standing for some time. You might need to decant them after 30 minutes. Do you see any change upon further standing? Pass light through each mixture and see which one

Competencies	*Contents*	*Suggested Activities*
		scatters light without separating on standing. What do you call this type of mixture? Students should come to the conclusion that both suspension and colloid are examples of heterogeneous mixtures whereas solution is a homogeneous mixture. Students should know that a solution has at least two pure components: solute and solvent. There could be more than one solute component in a solution but only one of the components is usually taken as the solvent. A substance dissolved in another substance is called solute and the substance in which the solute or solutes dissolve is the solvent Students should understand that solutions can assume the three state of matter and should know the general characteristic of solution formed by various combinations of gases liquids and solids.
• explain the different types of solutions • give examples for each types of solutions	**1.2 Types of solution** *(2 periods)* • Gaseous solutions	Gaseous solutions result when a gaseous solute dissolves in a gaseous solvent. Ask students to give an example for this. They may consider the components for solute or solvent. Nitrogen is taken to be the solvent and all other gases in air to be solutes. Ask them why.
	• Liquid solutions	Liquid solutions can have gaseous, liquid and solid solutes but the solvent is always liquid. Students should name some examples.
	• Solid solutions	Solid solutions have no restriction on the state of the solute but the solvent has to be solid. Since students may not be able to give examples of solid solutions readily as they did for the other two, you could take alloys as examples. An alloy is a solid solution of two or more metals or metals and non-metals. Tell students that they will learn more about alloying in Unit 5. Ask students if they know how jewelry gold is made. It is an alloy of cheaper

Competencies	*Contents*	*Suggested Activities*
		metals dissolved in gold, after both the solute and solvent are melted. The homogeneous mixture (solution) then solidifies upon cooling. You could arrange a visit to a nearby goldsmith for demonstration. This could also bring a positive change of attitude about the skilled workers in this sector of the economy.
• present a report on how jewelry gold is made to class after a visit to nearby goldsmith • explain how the "like dissolves like" rule depends on interparticle forces of interactions and predict relative solubilities;	**1.3 The solution process** *(5 periods)* • Liquid solutions and interparticle forces of attractions	In this section students will learn about the factors that control the solubility of substances in liquid solvents. Since water is the most important solvent in nature, the lessons need to be geared toward that. The key point here is that in order for substances to be appreciably soluble in each other, they must possess similar intermolecular attractive forces. Particles that attract each other very strongly tend to congregate and separate from those to which they are weakly attracted. Remember that when a solute particle is placed in solution it becomes solvated, which is surrounded by solvent molecules to which it is attracted. When the solvent is water, the term hydrated is used. You learned in Grade 11 chemistry how soaps and detergents work. This is a good excuse to revise about the cleansing actions of soaps and detergents. Soaps and detergents work on the principle of “like dissolves like.” Non-polar tails of soap anions dissolve in oil and grease globules. The polar heads keep the particles suspended in water so they can be washed away.
• define rate of dissolution; • define heat of solution, solvation energy and hydration energy. • apply the concept of heat of solution to the solution of ammonium nitrate crystal. • apply the concept of	• The rate of dissolution • Energy changes in the solution process - Heats of solution - Heats of hydration of ionic solid in water	The heat of solution is the energy absorbed or liberated when a solution is formed. For liquid solutions formed from a solvent A and solute B, an ideal solution results when the A-B attractions are the same as the A-A and B-B attractions. For an ideal solution, ΔH_{soln}

Competencies	*Contents*	*Suggested Activities*
heat of solution to the solution of sodium hydroxide crystal; • explain how heat of solution is influenced by the inter particle interaction forces.		= 0. When the A-B attractions are greater than the A-A and B-B attractions, $\Delta H_{soln} < 0$ and the solution process is exothermic. When the A-B attractions are weaker, $\Delta H_{soln} > 0$ and the solution process is endothermic. For solutions of solids in liquids the lattice energy (the energy required to separate the solute particles from a crystal) and hydration energy (or salvation energy- the energy released when the solute particle is placed into the solvent cage) must be considered.
• describe the distinctions among unsaturated, saturated and supersaturated solutions saturated solutions of sodium sulphate • prepare supersaturated solution of sodiumthiosulphate	**1.4 Solubility as an equilibrium process** *(4 periods)* • Saturated, unsaturated and supersaturated solution	Students have to appreciate because solutions are mixtures but not pure compounds they are not governed by the law of constant composition. Hence one can make solutions of different concentrations from the same solute and solvent. They can be unsaturated, saturated or even supersaturated. Students should be aware that a water-soluble substance dissolves in water until no more can dissolve at the given temperature. At the start the solution is unsaturated; it can dissolve more solute. Eventually the solution becomes saturated and can not dissolve any more solute. This behavior is also observed in solutions involving solvents other than water. • An **unsaturated solution** is a solution which can dissolve more solute at a given temperature. • A **saturated solution** is a solution which cannot dissolve any more solute at a given temperature Students should prepare unsaturated and saturated solution of sodium sulfate Students should also prepare supersaturated solution of sodiumthiosulphate. It can be prepared by adding sodiumthiosulpahte to hot water and letting it to cool slowly. Then add a small seed crystal to cause crystallization to occur.

Competencies	*Contents*	*Suggested Activities*
• explain the equilibrium nature of saturated solution • define solubility • describe the factors that affect solubility of substances • investigate the effect of temperature on solubility of sodium sulphate • conduct an experiment to determine solubility of table salt and sugar	• Effect of temperature on solubility	Students should know that a saturated solution is an equilibrium system in which the rate of dissolution is equal to the rate of crystallization. The solute is in equilibrium with its ions or molecules in solution. At a given temperature the concentration of the solute and its ions or molecules in a saturated solution are constant. Students should know that the solubility of a substance at a given temperature is the maximum quantity of substance, in moles or grams, which dissolves in a solvent to form 1liter of saturated solution at that temperature. The solubility of substances changes with temperature.
• state Henry's law • use Henry's law to calculate concentration of gaseous solute in a solution	• Effect of pressure on solubility of gases - Henery's Law	Ask students if they can draw parallel reasoning on factors that affect rates of dissolution with the overall factors that affect rates of chemical reactions. You may help them if they have forgotten what they learned in previous lessons. Among the many factors that determine the solubility, temperature should be given due consideration. A rise in temperature increases solubility if the dissolving of additional solute is endothermic. A fairly good rule of thumb is that the solubility of most solids and liquids in a solid solvent increases with increasing temperature. The solubilities of gases, however, almost always decrease with increasing temperature. Students should appreciate that the solubilities of almost all solid solutes increase with temperature whereas those of gases invariably decrease. Pressure has virtually no effect on the solubility of solids or liquids in liquid solvents. The solubilities of gases, however, are very markedly affected by pressure changes. Before you proceed ask students to use

Competencies	*Contents*	*Suggested Activities*
		Le Chatelier's principle to predict this. This is summarized in the form of Henry's law. Henry's law relates the concentration of a dissolved gas, Cg, to its partial pressure, Pg, over the solution. Cg = kgPg, where kg is Henry's law constant. Treat Henry's law only qualitatively.
• define concentration of a solution • define mass percentage, ppm and ppb of a solute in a solution • calculate the mass percentage, ppm and ppb of a solute in a solution from a given information • define mole fraction • calculate mole fraction of a solute and a solvent in a solution • define molarity • calculate molarity of a solution from a given information • define the terms equivalent mass, number of equivalents and normality • calculate normality of a solution from a given information • define molality • calculate molality of a solution from a given information • Inter convert various concentration expressions	**1.5 Ways of expressing concentrations of solution** *(7 periods)* • Mass percentage, ppm and ppb of solute • Mole fraction • Molarity • Normality • Molality • Conversion of concentration units	Students should be aware that the quantity of solute in a solution is a very important property of a solution. The quantity of solute per unit quantity of solution is called its **concentration.** Knowledge of it can be very useful in everyday life. For example, human beings estimate the quantity of solute in beverages, sugar, salt and other food solutions by tasting. Some farmers estimate the acid concentration in soils by tasting it to check whether the soil is suitable for planting crops. However, tasting is not an accurate and safe method of estimating the concentration of solute in solution. It doesn't allow the exact quantity of solute in solution to be known since different people have different tastes, apart from taste buds being unable to measure exact quantities. For example, many medicines in solution are prepared to specified concentrations of solute in solution to give correct dosage, otherwise they could be poisonous. Different forms of concentration arise due to the measurement of the quantities of solute and solvent in different units. Students should know the definitions of each concentration units. It is easy to confuse some of the units with similar names so students have to be cautious when using them. Students should know that there are many ways in which the concentration of a solution can be described. Students could describe the concentration of a solution using appropriate units such as : • Mass percentage, ppm and ppb of

Competencies	***Contents***	***Suggested Activities***
		solute • Mole fraction • Molarity • Normality and • Molality Let the Students calculate mass percentage, ppm, ppb, molarity, molality and normality of a solution from given information. Students should learn how to convert among the different units.
• prepare molar solutions of different substances • prepare normal solutions of different substances • prepare molal solutions of different substances • explain dilution process • calculate the volume or concentration changes during dilution of solution • prepare a dilute solution from concentrated solution	**1.6 Preparation of solutions** *(2 periods)*	Students should be aware that a common task in school, medical, industrial and other chemical laboratories is the preparation of solutions of known concentration. For aqueous solutions, distilled, demineralized or deionized water is used. Other solvents can also be used depending on the solution specified. Solutions are usually prepared from liquids or solids, and occasionally gases as solutes. Solutions in school science laboratories are prepared by teachers or laboratory technicians. Students should be able to calculate the quantities needed to prepare solutions. Students should know that a solution whose concentration is accurately known is called a **standard solution.** Students should prepare 1M, 1N and 1m solution of available chemical in a laboratory.
	• **Diluting solutions**	Students should be aware that stock solutions are prepared very concentrated. When we are to use them they usually need to be weakened or diluted. Students have to use the following formula to dilute solutions. $C_iV_i = C_fV_f$ where C_i is initial concentration, C_f is final concentration, V_i is initial volume and V_f is final volume. Ask students to prepare a 100ml,1M sucrose solution and then take a certain volume of it to prepare 100ml, 0.1 M solution by dilution.

Competencies	*Contents*	*Suggested Activities*
• use stoichiometerically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substances	**1.7 Solution Stoichiometry** *(3 periods)* • Mole - mass • Mole - volume • Mole - number of particles	Students should know that the principles of stoichiometry apply to reactants in solutions as well. Like in any other reaction, students should write balanced chemical equations that accurately describe the chemical reaction. Students should be able to appreciate that it is the particles dissolved in the solvent (assuming the solvent does not participate in the reaction) that take part in the reaction. If a solution is diluted by simply adding to it, the number of particle reactants remains the same. Students have to be aware of this when doing solution stoichiometry. Students could calculate amounts of reactants and products in a reaction of pure and dissolved substance.
• explain the relationship between reacting ions, spectator ions, precipitation and solubility • write net ionic equations	**1.8 Describing reactions in solution** *(1 period)* • Molecular equation • Ionic equation	Students should know that if all the ions in an ionic equation for a reaction do not cancel, then there is a net ionic equation and a reaction will occur. However, if all the ions do cancel, then there is no net ionic equation and no net reaction. When writing ionic equations, formulas for weak electrolytes are written in molecular form.
• define colligative properties • list the important properties of solvents that are affected by the formation of a solution • state Raoult's law • explain the change in the colligative properties of solvent when a certain solute is added • describe Vant Hoff's factor • calculate vapor pressure, the boiling point, and freezing	**1.9 Colligative properties of solutions** *(5 periods)* • Vapor pressure lowering • Boiling point elevation • Freezing point depression • Osmotic pressure	Keep the discussion qualitative so that students do not get bored by the arithmetic involved. You should of course be prepared to give additional work to able students. Students should appreciate that solutions have different physical properties from those of the pure solvents they are made of. Interestingly, these differences in properties depend not on the nature of the solute dissolved in the solvent but on the number of discrete particles per unit volume of the solution. Hence they are called colligative, because the word colligative means collective. The colligative properties of solutions are 1. vapor pressure lowering 2. boiling point elevation

Competencies	*Contents*	*Suggested Activities*
point of a solvent afte a certain amount of solute is added • define Osmosis and Osmotic pressure of a solution • calculate the osmotic pressure of a solution • compare and contrast change in colligative properties of electrolytic and non electrolytic solution		3. freezing point depression and 4. osmotic pressure Even though most of these effects are small, they have many practical applications, including some that are vital to biological systems. Ask students if equal masses of different solutes in the same amount of solvent will show the same vapor pressure, same boiling point, same freezing point and same osmotic pressure. You may tell students, in the temperate latitudes where winters are unbearably called, antifreeze such as ethylene glycol are added to radiators of cars. Ask for explanation. People often talk about bodies of cars imported used being damaged. Provoke discussion with colligative properties of solutions in mind. Introduce the formulas for all four colligative properties of solutions and make sure students have at least qualitative understanding of them. Make especial note of the similarity in form between the ideal gas equation and the van't Hoff's equation for osmotic pressure. Give some examples to show how osmotic pressure is used in biochemical laboratories to determine molecular masses of huge molecules like proteins and nucleic acids.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms mixture, homogenous and heterogeneous mixtures, solute, solvent, solution, rate of dissolution , heat of solution, solvation energy , hydration energy, solubility, concentration of a solution, mass percentage ,ppm and ppb of a solute in a solution, mole fraction, molarity , molality, equivalent mass, number of equivalents and normality; distinguish between homogenous and heterogeneous mixtures; explain the different types of solutions and give examples for each types of the solutions; explain how the "like dissolves like" rule depends on inter-particle forces of interactions and predict relative solubilities; discuss the factors that affect the rate of dissolution; apply the concept of heat of solution to the solution of ammonium nitrate and sodium hydroxide crystals; explain how heat of

solution is influenced by the inter particle interaction forces; describe the distinctions among unsaturated, saturated and supersaturated solutions; explain the equilibrium nature of saturated solution; describe the factors that affect solubility of substances; conduct an experiment to determine solubility of table salt and sugar; state Henry's law and use it to calculate
concentration of gaseous solute in a solution; calculate the mass percentage, ppm and ppb of a solute in a solution , mole fraction of a solute and a solvent in a solution, molarity of a solution, normality of a solution, and molality of a solution from a given information; prepare molar solutions, normal solutions and molal solutions of different substances; inter convert concentration units; explain dilution process and calculate the volume or concentration changes during dilution of solution; prepare a dilute solution from concentrated solution; use stoichiometerically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substances; explain the relationship between reacting ions, spectator ions, precipitation and solubility; write net ionic equations; define colligalive properties, Osmosis and Osmotic pressure of a solution; list the important properties of solvents that are affected by the formation of a solution; state Raoult's law; explain the change in the colligative properties of solvent when a certain solute is added; calculate vapor pressure boiling point, and freezing point of a solvent after a certain amount of solute is added, and the osmotic pressure of a solution; compare and contrast change in colligative properties of electrolytic and non electrolytic solution.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 2: Acid-Base Equilibria *(26 periods)*

Unit outcomes: Students will be able to:

- understand the Arrhenius, Bronsted-Lowery & Lewis concepts of acids & bases;
- understand the dissociation of water, weak monoprotic & polyprotic acids, & weak bases;
- know how to solve equilibrium problems involving concentration of reactants & products, k_a, k_b, P^H & P^{OH} ;
- understand the common ion effect, buffer solution, hydrolysis of salts , acid-base indicators and acid-base titrations;
- explain how buffering action affects our daily lives using examples;
- determine the equivalents of acid or base that are required to neutralize specific amount of acid or base;
- predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic or neutral;
- know how to solve problems involving concentration and P^H of acid-base titration;
- describe scientific enquiry skills along this unit: classifying, communicating, asking questions, applying concepts and making generalization.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • define acid by the Arrhenius concept • explain why proton exists bounded to water molecule , as H_3O^+, in all acid-base reactions; • give examples of Arrhenius acids • define base by the Arrhenius concept • give examples of Arrhenius bases;	**2. Acid-base Equilibria** **2.1 Acid-base concepts** *(5 periods)* • Arrhenius concept of acids and bases	It has to be clear from the outset that this unit deals quantitatively with the equilibria involving the self ionization of water and the dissociation of weak (not strong) acids and bases. These are important in any aqueous system, particularly biological ones where many important molecules behave as weak acids or bases. • Since many of the numerical problems in this unit require the application of some simple algebra, ample examples with detailed step-by-step procedures need to be given and students need to be warned not to rush and skip steps in the reasoning.
• define acid by the Bronsted-Lowry concept; • give examples of Bronsted-Lowry acids • define base by the Bronsted-Lowry concept • give examples of Bronsted-Lowry bases • explain what conjugate acids and conjugate bases are • identify the acid-base conjugate pairs from the given reaction; • write an equation for self-ionization of water	• Bronsted-Lowry concept of acids and bases - Conjugate acid-base pairs - Auto ionization of substances - Amphiprotic species	• Definitions of acids and bases according to Arrhenius, Bronsted-Lowry and Lewis should be discussed in groups by the students. Students should be able to see that they are not in contradiction to one another and hence each one of them can still be used; but that one is more comprehensive than the others. They should be able to see

Competencies	*Contents*	*Suggested Activities*
and ammonia • explain what is meant by amphiprotic species • give examples of reactions of amphiprotic species; • define acid by the Lewis concept • give examples of Lewis acids • define base by the Lewis concept • give examples of Lewis bases • describe the ionization of water	• Lewis concept of acids and bases	this for themselves during the discussion. Students should be aware that by Bronsted-Lowry definition, a great variety of chemical properties and chemical reactions can be correlated, including reactions that take place in solvents other than water or in solvent at all. Students should know some terms like conjugate acid-base pair, auto ionization and amphiprotic species under Bronsted-Lowry concepts. Students should be able to give examples of Arrhenius, Bronisted-Lowry and Lewis acids and identify conjugated acid – base pairs.
• drive the expression of ion product for water, K_W • explain the effect of temperature on Kw • explain why water is a weak electrolyte • use Kw to calculate $[H_3O^+]$ or $[OH^-]$in aqueous solution • define P^H • define P^{OH} • explain the relationship between P^H and P^{OH} • calculate P^H from $[H^+]$and $[H^+]$ from P^H • calculate P^{OH} from $[OH^-]$ and $[OH^-]$ from P^{OH} • write an expression for the percent ionization of weak acids or weak bases • calculate the percent dissociation of weak acids and bases • write the expression for the acid dissociation constant, K_a; • calculate K_a for an acid	**2.2 Ionic Equilibria of weak acids and bases** *(9 periods)* • Ionization of water - Ion product for water, K_W • Measures of the strength of acids and bases in aqueous solution; - H^+ ion concentration, $[H^+]$, and OH^- ion concentration , $[OH^-]$ - P^H and P^{OH} - percent ionization - Ionization (dissociation) constant * Acid ionization constant, K_a * Base ionization constant, K_b	Students should know that the most general theory of acids and bases is that of Lewis. This theory encompasses the Bronsted-Lowry theory, and it goes one step further by ascribing acid-base behavior to many reactions that do not involve the transfer of a proton. Students should appreciate one direction in which the science of chemistry matured, as narrow definitions of acids and bases progressively widened to encompass different species, physical states, solvent systems and reaction types. Students should know that equilibria involving the ionization of weak acids and weak bases in aqueous solution are very important in many fields of science. • Students have to know that in any solution in which water is the solvent, the equilibrium condition for the ionization of water $Kw = [H^+]\ [OH^-] = 1.0 \times 10^{-14}$ holds at 25 ^{0}C.

Competencies	***Contents***	***Suggested Activities***
from the concentration of a given solution and its P^H; • Calculate $[H^+]$ and P^H of an acidic solution from given values of K_a and initial concentration of the solution. • Write the expression for the base dissociation constant, K_b • Calculate K_b for a base from the concentration of the basic solution and its P^{OH}. • Calculate the $[OH^-]$ and POH of a basic solution from a given values of K_b and initial concentration of the solution.		They should also be able to calculate pH/pOH given $[H^+]$ or $[OH^-]$and $[H^+]$ or $[OH^-]$ given pH/pOH . Students should understand that the strengths of acids and bases can be described in many ways. Some of the ways are: • H^+ and OH^- ions concentration • P^H and P^{OH} • Percent dissociation • K_a and K_b Here K_a and K_b are treated more. Students should know that distilled or deionized water is an extremely weak electrolyte. It conducts electricity very weakly and any conduction by it can only be detected by very sensitive galvanometers. Extremely low concentrations of hydrogen, H^+ and hydroxide, OH^- are in equilibrium with the undissociated water molecules, H_2O. Students could be asked to derive ion product for water, K_w, $\mathbf{K_w = [H^+]\,[OH^-]}$ Students should be aware that in any aqueous system, pure or not, the ionic product is always constant at a constant temperature. It doesn't matter where the hydrogen (H^+) and hydroxide (OH^-) ions come from. Students should appreciate that an aqueous solution can be acidic, neutral or basic. A solution is described as neutral when the hydrogen ion concentration is equal to the hydroxide ion concentration, $[H^+] = [OH^-]$. If the concentration of hydrogen ions is above 1×10^{-7} mol/l the solution is said to be acidic. If the

Competencies	*Contents*	*Suggested Activities*
		concentration of hydrogen ions is below 1 X 10^{-7} mol/l (i.e. the concentration of hydroxide ions is above 1 X 10^{-7} mol/l) the solution is said to be alkaline or basic. Students should be aware that P^H scale was introduced by Danish biochemist S. P. Sorensen in 1909. Chemists use the P^H scale to measure acidity or alkalinity of solutions. $\mathbf{P^H = - log [H^+]}$ $\mathbf{P^{OH} = - log [OH^-]}$ Students should know that the magnitude of acid ionization constant, K_a and the base acid ionization constant, K_b are measures of the acid strength of an acid and the base strength of a base respectively. • An acid with a higher value of K_a is a stronger acid than the acid with a lower value of K_a • An acid with a higher value of K_a dissociates more than the acid with a lower value of K_a • A base with a higher value of K_b is a stronger base than the base with a lower value of K_b • A base with a higher value of K_b dissociates more than the base with a lower value of K_b Students could exercise to calculate K_a and K_b of an acid and a base from given solutions and their P^H and P^{OH} respectively.
• define common ion effect; • explain the importance of common ion effect; • define a buffer solution; • give some common	**2.3 Common ion effect and Buffer Solution** *(4 periods)* • Common ion effect	Common ion effect could be given to the students as a group work to see if they have understood the Le Chatelier's principle discussed in grade 11 Chemistry class. Effort has to be made to see that the students understand this topic

Competencies	*Contents*	*Suggested Activities*
examples buffer systems; • explain the action of buffer solution and it's importance in chemical processes • calculate the P^H of a given buffer solution • demonstrate the buffer action of CH_3COOH/CH_3COONa; • define hydrolysis;	• Buffer solutions	because it is a test case for many related topics. Students should know that buffer solutions have wide applications especially in biological systems. However since the concept is a formidable one, the discussion has to focus on applications rather than numerical problems. A buffer is a mixture of a weak acid and a weak base. It can be prepared by mixing a weak acid with one of its salts, or a weak base with one of its salts. A solution resists pH change in the presence of a buffer when a small amount of an acid or base is added, which otherwise would suffer pH change. Give your students how this is useful in biological systems. One of the first things your students learn in this section is that there is an inverse relationship between the strength of a BrØnsted acid and the strength of its conjugate base. Strong acids have very weak conjugate bases; strong bases have very weak conjugate acids. Students should demonstrate the buffer action of CH_3COOH/CH_3COONa.
• explain why a salt of weak acid and strong base gives a basic solution • explain why a salt of strong acid and weak base gives an acidic solution • explain why salts of weak acids and weak bases give acidic, basic or neutral solution	**2.4 Hydrolysis of Salts** *(2 periods)* • Salts of weak acids and strong bases: anion hydrolysis • Salts of strong acids and weak bases: cation hydrolysis	Students should know that solutions of salts contain two ions. The cation of the salt can potentially serve as an acid; the anion as a base. Whether a solution of a salt is acidic, basic, or neutral depends on the strengths of these acids and bases. In determining what happens in solution of a salt, students have to examine the salt to determine what its 'parents' are. For example, sodium chloride, NaCl has as its parents the acid HCl and the base NaOH. Each of these are strong, so we categorize NaCl as the salt

Competencies	*Contents*	*Suggested Activities*
• define acid-base indicators • write some examples of acid-base indicators • suggest suitable indicator for a given acid-base titration	• Salts of weak acids and week bases: cation and anion hydrolysis	of a strong acid and a strong base. On the other hand, sodium nitrite, $NaNO_2$, is the salt of a weak acid, HNO_2 and a strong base, NaOH. Students could discuss hydrolysis of each type of salt. They may raise the following points. **Salts of strong acids and strong bases**. Solutions of these salts are neutral because the ions are very weak acids and bases. Hence these ions don't hydrolyze. Examples are NaCl and KNO_3. Ask students to give more examples.
• explain the equivalents of acids and bases • calculate the normality of a given acidic or basic solution • define acid-base titration • define end point • define equivalence point • distinguish between end point and equivalent point;		**Salts of weak acids and strong bases**. Solutions of theses salts are basic because the anion of the weak acid is a moderately strong base . Examples are $NaC_2H_3O_2$ and $NaNO_2$. In general, the anion of the salt reacts as follows: $X^- + H_2O \leftrightarrow HX + OH^-$. Since a component of the salt is reacting with water, the phenomenon has come to be known as hydrolysis. **Salts of strong acids and weak bases.** Solutions of theses salts are acidic because the cation of the weak base is a moderately strong acid. Examples are NH_4Cl and N_2H_5Cl. **Salts of weak acids and weak bases**. In these solutions there is both a strong acid (the cation of the weak base) and a strong base (the anion of the weak acid). Whether the solution of such a salt is acidic, basic, or neutral depends on the relative strengths of the acidic cation and basic anion. If the acid is stronger than the base, the solution is acidic, and vice versa. If they are of equal strengths, the solution is neutral. This topic should be treated lightly and only qualitatively.

Competencies	***Contents***	***Suggested Activities***
• discuss the different types of titration curves	**2.5 Acid-base indicators and Titrations** *(6 periods)* • Acid-base indicators • Equivalents of acids and bases - Number of equivalents - Normality • Acid-base titrations • The equivalent point and the end point • Acid-base titration curves - Titration of a strong acid with a strong base	Students should understand that acid-base indicators are organic molecules whose colors depend on the acidity or basicity of the solution in which they are dissolved. This needs to be discussed from equilibrium point of view. pH meters can also be used to measure $[H^+]$ by means of two electrodes immersed in a test solution. This should not be discussed in detail but showing how to use the apparatus would be beneficial The addition of a common ion to a solution containing a salt in equilibrium with its ions decreases the solubility of the salt. This is in accordance to Le Chatelier's principle applicable to systems at equilibrium. Acid-base indicators are organic molecules that indicate whether a solution is acidic, basic or neutral. An acid-base indicator is a weak acid or weak base with its corresponding conjugate pair. The molecular form of the indicator, HIn, has one color and the ionic form In^-, has a different color. The color that is observed in a solution of the indicator is controlled by the ratio of [HIn] to [In], which is determined in turn by the H^+ concentration in the solution. An indicator can be used to signal the completion of a reaction during a titration by changing color. Since students find it difficult to understand how the equivalent mass is different from the molecular mass of a substance analogies could be used. An equivalent of an acid is the mass of the acid that releases one mole of H^+ in solution. Likewise an equivalent of a base is the mass

Competencies	*Contents*	*Suggested Activities*
	- Titration using a weak acid and a strong base - Titration of a weak base with a strong acid	of the base that produces one mole of OH^- or accepts one mole of H^+. Students should be able to make similar definitions for reductants and oxidants in redox reactions. Students should appreciate how the definitions relate to the various definitions of acids and bases. Students should be given numerous examples and problems to acquaint themselves with the difference between equivalent mass and molecular mass of a substance. Ask students to define molarity and normality and see if they know that the difference lies in noticing the differences between equivalent mass and molecular mass. Students could exercise to calculate the normality of a given acidic or basic solution. • Students should not confuse between titration, a method of determining how much acid or base is present in a solution and pH, which is a measure of acidity or basicity of a solution. It might be helpful to give examples like the following one. A 0.1 M solutions of HCl in water and a 0.1 M solution acetic acid have pH's 1 and 2.9, respectively, that is they differ in acidity. On the other hand the total concentration of acid is the same and each of these solutions neutralizes the same amount of NaOH solution. • Students usually think at the equivalence point of a titration the pH is always 7. They need to see that this is true only if the titration is between a strong acid and strong base. They will have to relate this to hydrolysis of anions or cations. Acids are titrated to the point of

Competencies	*Contents*	*Suggested Activities*
		neutralization by adding equivalent amount of a base. An indicator is employed to determine the end point. At the endpoint or equivalence point, equal numbers of equivalents of acids and bases will have been added. Titration is an analytical procedure in which a solution, generally of known concentration, is added gradually from a burette to another solution where the solutes react. This is continued until the completion of the reaction is signaled by an indicator which was added at the beginning of the titration. Acid base titration curves are drawn by plotting points corresponding to the pH of the solution and equivalents of the titrant added. Students could think that at the equivalence point the pH of the solution is 7. This may not be true always. This is true if hydrolysis of the salt formed from the acid base reaction is nonexistent. Salts of strong acid and strong base are neutral. Salts of weak acids and strong base are basic and salts of strong acids and weak base are acidic. A way to remember this is to consider hybrids of characteristics of different strengths: the stronger one dominates. What is more important is for the students to be able to explain this from hydrolysis of salts point of view.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define acids and bases by the Arrhenius, Bronsted-Lowry and Lewis concepts, and give examples for each of them; explain why proton exists bounded to water molecule , as H_3O^+, in all acid-base reactions; explain what conjugate acids and conjugate bases are and identify the acid-base conjugate pairs from the given reaction; write an equation for self-ionization of water and ammonia; explain what is meant by amphiprotic species and give examples of their reactions; describe the ionization of water and explain why water is a weak electrolyte; drive the expression of ion product for water, K_W , and use it to calculate $[H_3O^+]$ or $[OH^-]$ in aqueous solution; define P^H and P^{OH} , and explain the relationship between them; calculate P^H from $[H^+]$ and $[H^+]$ from P^H; calculate P^{OH} from $[OH^-]$ and $[OH^-]$ from P^{OH} ; write an expression for the percent ionization of weak acids or weak bases and calculate them; write the expression for the acid dissociation constant, $K_{a,}$ and the base dissociation constant, $K_{b;}$ calculate K_a for an acid from the concentration of a given solution and its P^H , and K_b for a base from the concentration of the basic solution and its P^{OH}; Calculate the $[H^+]$ and P^H of an acidic solution from given values of K_a and initial concentration of the solution , and the $[OH^-]$ and P^{OH} of a basic solution from a given values of K_b and initial concentration of the solution; define common ion effect and explain the importance of common ion effect; define buffer solution , give some common examples buffer systems, and explain the action of buffer solution and it's importance in chemical processes; calculate the P^H of a given buffer solution; prepare a simple buffer solution and test its buffering capacity; define hydrolysis; explain why a salt of weak acid and strong base gives a basic solution, a salt of strong acid and weak base gives an acidic solution, and a salt of weak acid and weak base gives acidic, basic or neutral solution; define acid-base indicators and write some examples of them; suggest suitable indicator for a given acid-base titration; explain the equivalents of acids and bases; calculate the normality of a given acidic or

basic solution; define acid-base titration, end point and equivalence point; distinguish between end point and equivalent point; discuss the different types of titration curves;

Students above minimum requirement level

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Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 3: Introduction to Chemical Thermodynamics *(12 periods)*

Unit outcomes: Students will be able to:

- understand the terms systems (open, closed & isolated), spontaneous process, functions(state & path), properties (intensive & extensive), internal energy, heat & work;
- understand the first and second laws of thermodynamics;
- compare the energy changes observed when chemical bonds are formed & broken, & relate these changes to endothermic & exothermic reactions;
- identify ways in which the terms reactants, products & the heat are combined to form thermo chemical equations representing endothermic & exothermic chemical changes;
- understand concepts such as enthalpy change, entropy & entropy changes & free energy change;
- determine the spontaneity of a given reaction;
- describe scientific enquiry skills along this unit: observing. Communicating, measuring, applying concept and asking questions.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • define chemical thermodynamics; • define systems; • give examples of systems; • classify system into open, closed and isolated systems' • explain open closed and isolated system • define state function and path function • distinguish between intensive and extensive properties, • define spontaneous and non-spontaneous process • explain internal energy in relation to the concepts of thermodynamics; • explain heat in relation to the concepts of thermodynamics, ▪ explain work in relation to the concepts of thermodynamics, • state first law of thermodynamics,	**3. Introduction to Chemical Thermodynamics** **3.1 Common thermodynamic terms** *(2 periods)* • Systems, Surroundings, Functions, Properties and Processes	Students should know two factors that control whether the products of a chemical reaction will form are kinetics and thermodynamics. Thermodynamics controls the feasibility of a reaction in the sense that it determines whether a reaction is possible and how much products can be formed. It deals with energy changes and is applied to physical as well as chemical changes. Students find it easier to comprehend the principles of chemical thermodynamics if the principles are developed using physical systems followed by the extension of the principles to chemical systems. • Common thermodynamic terms need to be elaborated with examples and question & answer methodology. Students should not move on before mastering these essential terms. Ask students to read and understand the meanings of the following terms and write the definitions in their own words. It would be good if they try to explain them verbally to their class mates. Thermodynamics System Surroundings Isothermal Adiabatic State

Competencies	*Contents*	*Suggested Activities*
- explain first law of thermodynamics, - calculate change in internal energy of a system from a given information,		Equation of state Heat capacity Specific heat State function State variable
- define enthalpy change, ΔH - derive an expression for the enthalpy change of an ideal gas based on first law of thermodynamics, - calculate enthalpy change for ideal gases from given information,	**3.2 First Law of thermodynamics and some thermodynamic quantities** *(3 periods)* - Internal energy (E) - Heat (q) - Work (W) - First Law of Thermodynamics	- First law of thermodynamics has to be kept to its bare minimum; however, the various forms of its definition need to be discussed in groups for the purpose of clarity. During a chemical change the energy gained by the system, (the reaction mixture inside the vessel) is equal to the energy lost by the surroundings to the system and vice versa. In other words energy cannot simply appear or disappear- it is neither created nor destroyed. The total energy of the universe is constant. These are different forms of the first law of thermodynamics.
	3.3 Thermo chemistry *(4 periods)* - Heats of reaction (Enthalpy changes)	Ask students if a glass of water at room temperature will freeze. If it does freeze, it has lost energy and the surrounding will gain equal amount of energy. This does not happen not because the first law is not obeyed. They should consider the enthalpy change and see if it makes sense from what they see. How about an ice cube at room temperature melting? Is its enthalpy increasing or decreasing? Encourage them to argue among themselves by citing similar examples. Are they convinced that decrease in enthalpy of a system by itself alone is not enough to be taken as an indicator of spontaneity? When a system undergoes a series of changes that ultimately brings it back to its

Competencies	***Contents***	***Suggested Activities***
- define standard state; - define standard molar enthalpy of combustion, ΔH^0 comb; - describe how heats of combustion can be used to estimate the energy available from foods - define standard molar enthalpy of formation, ΔH^0 form; - define standard molar enthalpy of neutralization, ΔH^0 neutralization; - carry out an activity to measurer standard molar enthalpy of neutralization; - state Hess's law; - apply Hess's law to solve problems on enthalpy changes of chemical reactions;	- Standard states	original state, the net energy change for the system is zero. This is a formal statement of the law of conservation of energy, and is the basis of Hess's law. Often stated mathematically as $\Delta E = q - w$, where ΔE is the change in the internal energy, q is the heat added to the system, and w is the work done by the system. Internal energy, E is the total kinetic and potential energy of a system. At this level do not try to show the subtle differences between enthalpy change and internal energy change. After all even in reactions where gases are involved the difference is usually less than 1%. It might be sufficient to show that $\Delta E = q_v$ (heat absorbed or evolved at constant volume) and $\Delta H = q_p$ (heat absorbed or evolved at constant pressure) Students should be aware that the standard states for thermodynamic properties are not the same as the standard states used to describe gases. Standard conditions are chosen to be 298 K and 1 atm pressure. A substance in its natural state under these conditions is said to be in its standard state. Standard states are indicated by a superscript zero.
- explain bond energy' - calculate the bond energies of substances in a given chemical reaction' - calculate the standard enthalpy changes of a reaction from given enthalpy changes of reactants and products;		Students should determine the enthalpies of neutralization of some acids and bases.
	- Hess's Law	Students should appreciate that Hess's law of heat summation is an

Competencies	*Contents*	*Suggested Activities*
		important tool in writing a thermo chemical equation. A chemical change written to show the energy change that takes place is called a thermo chemical equation. Hess's law says, in effect, that the ΔH for some net reaction is the sum of all of the ΔH's for steps along the way. When thermo chemical equations are added together to obtain some final equation, the ΔH for the final equation is the sum of the ΔH's of the thermo chemical equations that were combined. To help students have a clearer understanding of the concept, assignments should be given.
	• Bond Energies	Students should know that thermodynamic data can be used to obtain information about strengths of bonds. Likewise bond energies can be utilized to obtain estimate of the heat of formation of a compound. Students should be able to do numerical problems that will help them understand the concept better.
• explain entropy and entropy change; • calculate the entropy change, from the given standard entropies of substances, • state the second law of thermodynamics; • explain the second law of thermodynamics; • calculate the entropy changes from the given enthalpy change of the system and absolute temperature; • explain free energy and free energy change; • calculate the standard free energy change, ΔG^0, from given standard free energies	**3.4 Entropy and Second Law of Thermodynamics** *(3 periods)* • Entropy(s) and Spontaneous process	Students should appreciate that the second law of thermodynamics is the most applied tool in chemical thermodynamics; it is the criterion of spontaneity. Since spontaneity is easily confused with speed by students, enough time should be allotted to explain what it exactly means. The $\Delta G = \Delta H - T\Delta S$ criterion has to be explored using a table in group discussions. This has to be limited, though, to qualitative discussions. A spontaneous change is one that occurs without **continued outside help.** It may need a push to get it started (the mixture of H_2 and O_2, for instance) but once started, it

Competencies	***Contents***	***Suggested Activities***
of reactants and products • describe the relationship between standard free energy change, ΔG^o, standard enthalpy change, ΔH^0, and standard entropy change, ΔS^o, of a reaction, • determine the spontaneity of a given reaction.	• Second law of Thermodynamics - Free energy - Criterion for spontaneous process: ΔS, ΔG, ΔH	continues on its own. Students should understand that there are two thermodynamic quantities that determine whether events are spontaneous. Spontaneity is favored if there is an energy decrease (if ΔH is negative, corresponding to an exothermic change). It is also favored if there is an increase in the degree of randomness of a system. The thermodynamic quantity related to randomness is entropy, S. Spontaneity is favored by an entropy increase (if ΔS is positive). If both changes in enthalpy and entropy are favorable, i.e ΔH is negative and ΔS is positive then definitely the process is favored. However, if ΔH is positive and ΔS is negative the process is not favored. Students have to be aware that chemical reactions that are not favored can take place if energy from outside is continually supplied. A corollary to this is that reactions that are thermodynamically allowed may not take place until the initial push is provided. The initial push could be small or big which was discussed in Grade 11, under the topic chemical kinetics. Students should know that the free energy change, ΔG is equal to the maximum work that can be obtained from a spontaneous change. It is also an indicator for the spontaneity of a process; if ΔH is < 0, spontaneous if ΔH is > 0, non-spontaneous and if ΔH is $= 0$, at equilibrium.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define chemical thermodynamics; define systems and give examples of systems; classify systems into open, closed and isolated systems and explain them; define state function and path function; distinguish between intensive and extensive properties; define spontaneous and non-spontaneous processes; explain internal energy, in relation to the concepts of thermodynamics; explain heat, internal energy and work in relation to the concepts of thermodynamics; state and explain first law of thermodynamics; calculate change in internal energy of a system from a given information; define enthalpy change, ΔH, derive an expression for the enthalpy change of an ideal gas based on first law of thermodynamics and calculate enthalpy change for ideal gases from given information; define standard state, standard molar enthalpy of combustion, ΔH^0 comb, standard molar enthalpy of formation, ΔH^0 form, and standard molar enthalpy of neutralization, ΔH^0 neutralization; describe how heats of combustion can be used to estimate the energy available from foods; carry out an activity to measurer standard molar enthalpy of neutralization; state Hess's law and apply it to solve problems on enthalpy changes of chemical reactions; explain bond energy and calculate the bond energies of substances in a given chemical reaction; calculate the standard enthalpy changes of a reaction from given enthalpy changes of reactants and products; explain entropy and entropy change, and calculate the entropy change from the given standard entropies of substances; state and explain the second law of thermodynamics; calculate the entropy changes from the given enthalpy change of the system and absolute temperature; explain free energy and free energy change; calculate the standard free energy change, ΔG^0, from given standard free energies of reactants and products; describe the relationship between standard free energy change, ΔG^o, standard enthalpy change, ΔH^0, and standard entropy change, ΔS^o, of a reaction; determine the spontaneity of a given reaction.

Students above minimum requirement level

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Students below minimum requirement level

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Unit 4: Electrochemistry (*21periods)*

Unit outcomes: Students will be able to:

- understand the fundamental concepts related to oxidation- reduction reaction;
- know the application of redox reactions in production of new substances and energy;
- demonstrate an understanding of fundamental concepts related to the interconversion of chemical & electrical energy;
- understand the difference between metallic conduction & electrolytic conduction;
- Identify & describe the functioning of the components of electrolytic & galvanic cells;
- understand the difference between electrolytic & galvanic cells;
- know how to solve problems based on Faraday's law;
- measure through experimentation the mass of metal deposited by electroplating (e.g.:- copper from copper(II) sulphate), & apply Faraday's law to relate the mass of metal deposited to the amount of charge passed;
- predict the spontaneity of redox reactions & overall cell potentials by studying a table of half-cell reduction potentials;
- determine the emf of an electrochemical cell experimentally or from given data;
- explain the application of electrochemistry in our daily lives & in industry;
- explain corrosion as an electrochemical process, & describe corrosion- inhibiting techniques (e.g. painting, galvanizing, cathodic protection);
- describe examples of common galvanic cells & evaluate their environmental & social impact;
- describe scientific enquiry skills along this unit: observing, classifying, comparing & contrasting, communicating, asking question, measuring, relating cause & effect and problem solving.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • define redox reaction • define oxidation in terms of electron transfer and change in oxidation number. • define reduction in terms of electron transfer and change in oxidation number. • describe the oxidizing and reducing agents • identify the species that are oxidized and reduced in a given redox reaction and determine the oxidizing and reducing agents • balance a given redox reaction using oxidation number	**4. *Electrochemistry*** **4.1. Reduction-oxidation reactions** *(5 periods)* • Oxidation • Reduction • Balancing redox reactions - Change in Oxidation number method - Ion-electron method	Since Redox Reactions play important role in all areas of chemistry especially Biological sciences, students should be encouraged to think of chemical reactions of relevance to plant and animal physiology. Remind students about rules to assign oxidation numbers. Let them understand the terminologies like redox reaction, oxidation ,reduction half-reaction, oxidizing agent and reducing agent. Ask students to identify the oxidized and reduced species and, the oxidizing and reducing agents. To help students the concept that an oxidizing agent it self reduced, you can discuss the action of a bleaching agent on clothing in a washing machine. A bleaching agent, which causes a whitening to occur, does not itself become whiter. An oxidizing agent is not oxidized.

Competencies	*Contents*	*Suggested Activities*
method • balance a given redox reaction using ion-electron method • Recall metallic conductivity, electrolytic conductivity and electrolysis of molten electrolytes. • draw labeled diagrams of an electrolytic cell; • define preferential discharge • explain factors that affect preferential discharge. • describe the effect of nature of the ions on the electrolysis of aqueous solutions of dil. H_2SO_4 and NaOH. • describe the effect of concentration of the ions on the electrolysis of dilute and concentrated solutions of NaCl • describe the effects of types of electrodes on the electrolysis of $CuSO_4$ solution using Ni electrodes and Cu electrodes. • write electrode half reactions and overall reactions for the electrolysis of dil. H_2SO_4, NaOH, dil	**4.2 Electrolysis of Aqueous Solutions** *(6 periods)* • Electrolytic cell • Preferential discharge • Electrolysis of some selected aqueous solutions	Students should be guided step by step on how to use the Change in Oxidation Number method and Ion-Electron method to balance equations where the simpler methods do not help. It is important to point out to students that oxidation numbers used in balancing redox equations are arbitrary values obtained by using arbitrary rules. Students should be warned against thinking that the oxidation number represents the real ion charge in the molecule. Students could discuss the metallic and electrolytic conductivities and electrolysis of molten electrolytes. All chemical changes, including electrochemical ones, involve energy changes. The main focus of this unit is to show how redox reactions can be caused to occur by the action of electricity, and the way electricity can be obtained from redox reactions that occur spontaneously. Students should appreciate the processes that take place at the cathode and anode in an electrolytic cell. The terminologies- cathode, cation, anode and anion are easily confused by students. Students should discuss in groups until they have them right. Students should learn from examples to be given how the net reaction in the electrolysis of aqueous solutions is controlled by which redox reactions occur most easily. Students should be asked to draw and label electrolytic cells. They should also show at which electrodes reduction and oxidation reactions take place. Students should know that the electrode can discharge only one type of species in preference to the other at a time Students should be able to confirm that electrons lost during oxidation half

Competencies	*Contents*	*Suggested Activities*
NaCl, conc. NaCl and $CuSO_4$ solutions • distinguish between molten electrolytes and aqueous electrolytic solutions,		reaction are equal to electrons gained during reduction. The net reaction in a cell is the sum of cathode (reduction) and anode (oxidation) half reactions. Here the connection to the ion-electron method of balancing equations. Students should be assisted to conduct experiments on electrolysis of aqueous solutions and confirm the effects of different factors on preferential discharge of ions. They should write the anode, cathode and overall reactions name the product for each types of activities.
• State Faraday's first law of electrolysis. • write the mathematical expression for the Faraday's first law of electrolysis, • do calculations related to Faraday's first law of electrolysis, • state Faraday's second law of electrolysis. • write the mathematical expression for the Faraday's second law of electrolysis.	**4.3 Quantitative Aspects of Electrolysis** *(3 periods)* • Faraday's First Law of Electrolysis • Faraday's Second Law of Electrolysis	Ask students to tell the difference between the electrolysis of molten electrolytic and aqueous electrolytic solutions. Students should know that electrolytic reactions are no exception to the general quantitative relationship of the amount of energy involved and the amount of chemical produced. Students should be asked to formulate the two Faraday's laws of electrolysis. Quantitative and qualitative significance of the laws have to be explored from numerous qualitative and quantitative problems. Students should be asked to practice calculations involving Faraday's laws of electrolysis as class work and assignment.
• do calculations related to Faraday's second law of electrolysis. • mention industrial applications of electrochemistry, • explain electroplating and electro refining, • demonstrate copper refining, • explain how electrolysis is used in the production of some metals, nonmetals and	**4.4 Industrial Application of Electrolysis** *(2 periods)* • Electroplating and Electro refining • Extraction of metals • Manufacturer of Non-metals and some compounds	Electrolysis has wide applications. Students should be asked to list down some of them in class. They should also be asked to explain to class how they are related to electrolysis. Let the students conduct an experiment on refining of copper. When they discuss electroplating they should know which electrode is the metal to be plated and vice versa. Ask students which elements are extracted from their ores only through electrolysis and why.

Competencies	*Contents*	*Suggested Activities*
compounds, • define voltaic cell and salt bridge • draw and label Zn - Cu voltaic cell, • define electrode potential and cell potential, • construct Zn - Cu voltaic cell	**4.5 Voltaic Cells** *(5 periods)*	Students should know that voltaic cells have important roles to play in society. The cells are divided into primary and secondary. Students have to be careful when handling such devices. Ask the students to write chemical equations for the reactions they think take place. They should be able to identify the various components of the cells. Ask students to draw diagrams of the various types of voltaic cells.
• measure the cell potential of Zn - Cu cell using voltmeter • explain how standard electrode potential is measured • describe the reactivity of a metal from its position in the activity series or electro motive series	• Construction of Zn - Cu voltaic cell • Cell diagram (Notation for Voltaic cell) • Electromotive Force • Measuring standard electrode potentials	Students should construct a simple galvanic cell (Daniel's cell). They should identify the parts and measure the cell potential. They should also measure potentials of half-cell using standard half-cells. Finally students can make their own EMS of given elements.
• calculate cell potential • decide whether a given redox reaction is spontaneous or not. • explain the effect of concentration on cell potential • Mention the different types of voltaic cells • give examples of each type of voltaic cell • describe how Hydrogen-Oxygen Fuel cell operates • distinguish between primary and secondary cells • identify the cathode, anode, and the electrolyte of a given voltaic cell, • compare and	• Calculation of cell potential • Dependence of emf on concentration. • Types of voltaic cells, • Corrosion of metals.	Students should exercise calculating the cell potential and decide the spontaneity of the reactions. Assign students to collect and bring used and disposed batteries and ask them to identify their types. Students have to be careful when handling such devices. Students should understand how hydrogen-oxygen fuel cell operates. Students should able to tell the differences between primary, secondary and fuel cells. Help students design experiments that will help them identify the factors responsible for corrosion f metals. Can the students see that this is a spontaneous reaction. Would we have to worry if it was a non-spontaneous reaction?

Competencies	*Contents*	*Suggested Activities*
contrast electrolytic and voltaic cells • explain metallic corrosion in terms of redox reaction • explain the negative effects of corrosion, • explain the different methods of prevention of corrosion		

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define redox reaction, oxidation and reduction; identify the species that are oxidized and reduced in a given redox reaction, and describe the oxidizing and reducing agents; balance a given redox reaction using oxidation number method and ion-electron method; distinguish between molten electrolytes and aqueous electrolytic solutions; draw labeled diagrams of an electrolytic cell; define preferential discharge and explain factors that affect it; describe the reactivity of a metal from its position in the activity series or electro motive series; describe the effect of nature of the ions on the electrolysis of aqueous solutions of dil. H_2SO_4 and NaOH; describe the effect of concentration of the ions on the electrolysis of dilute and concentrated solutions of NaCl; describe the effects of types of electrodes on the electrolysis of $CuSO_4$ solution using Ni electrodes and Cu electrodes; write electrode half reactions and overall reactions for the electrolysis of dil. H_2SO_4, NaOH, dil NaCl, conc. NaCl and $CuSO_4$ solutions; state Faraday's first law of electrolysis , write its mathematical expression and do calculations related to it; state Faraday's second law of electrolysis, write its mathematical expression and do calculations related to it; mention industrial applications of electrochemistry; explain electroplating and electro refining; explain how electrolysis is used in the production of some metals, nonmetals and compounds; define voltaic cell , salt bridge , electrode potential and cell potential; draw , label and construct Zn - Cu voltaic cell; measure the cell potential of Zn - Cu cell using voltmeter; explain how standard electrode potential is measured and calculate it from the given information; decide whether a given redox reaction is spontaneous or not; explain the effect of concentration on cell potential; mention different types of voltaic cells and give examples of each type of them; distinguish between primary and secondary cells; describe how Hydrogen-Oxygen Fuel cell operates; identify the cathode, anode, and the electrolyte of a given voltaic cell; compare and contrast electrolytic and voltaic cells; explain metallic corrosion in terms of redox reaction and the negative effects of corrosion; explain the different methods of prevention of corrosion.

Students above minimum requirement level

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Students below minimum requirement level

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Unit 5: Some Elements in Nature and Industry *(17 periods)*

***Unit out comes*:** Students will be able to:

- describe occurrence and abundance of the elements in nature;
- explain how carbon, nitrogen and phosphorus cycle in nature;
- understand the metallurgical processes;
- understand the occurrence, extraction & chemical properties of Sodium, Calcium, Tin, Lead, Zinc & Chromium;
- explain the steps in industrial production of ammonia, nitric acid, sulfuric acid and Diammonium phosphate.
- describe scientific enquiry skills along this unit: classifying, communicating, asking questions, relating cause & effect and making generalization.

Competencies	***Contents***	***Suggested Activities***
Students will be able to: • define the terms mineral and ore	***5. Some Elements in Nature and Industry*** **5.1 Some Elements in Nature** *(5 periods)*	This unit is believed to be handled with ease if portions are assigned to students for class presentation. The students need to be encouraged to bring metal pieces and non-metals from their neighborhoods for viewing.
• describe the forms of occurrences of metals • discuss the distribution and relative amounts of the elements in the earth's crust	**5.1.1 Occurrence of elements** • Sources of elements • Abundance of elements	There are 92 elements found in nature. All natural elements are made out of them. Students should be aware that elements can be found either pure or as part of a compound. The elements, in their pure form, are found in one of the three states at ambient conditions of temperature and/or pressure. Most non-metallic elements are gases; whereas most metallic elements are solids. Variety of techniques can be applied to separate elements in their pure form from their sources based on differences in physical properties. Some elements are not found abundantly in pure form; therefore they are extracted from compounds where they are abundant utilizing physical and chemical methods. Most metallic elements, however, are found in ores, substances that contain desirable constituents in concentrations large enough to make their recovery economically feasible.
• define the term fixation • discuss the carbon cycle, the nitrogen cycle and the phosphorus cycle	**5.1.2 The recycling of elements in nature** • The carbon cycle • The nitrogen cycle	

Competencies	*Contents*	*Suggested Activities*
	- The phosphorus cycle	The financial costs of mining, separating and purifying an element are considered before a process is chosen for implementation. Students already know matter is conserved; this of course does not preclude elements from undergoing chemical and physical changes that affect the amount of a certain element in a given form and at a given location without affecting the total sum in the different layers of the crust: lithosphere, hydrosphere and atmosphere. Students should appreciate that carbon is found in all layers of the earth's crust and living things. It is found as CO_2 (g) in the air which can be removed in the form of glucose during photosynthesis. It then becomes part of the plants and animals that feed on plants. The dissolved CO_2 in the ocean becomes part of the various carbonates. Ask students what impact afforestation, deforestation and building plenty of cement factories will have on global warming and why. Describe the role of photosynthesis, respiration, and decay in the carbon cycle. Students should appreciate that nitrogen is 80% of the air in the atmosphere but is not usable in this form by living things. Students should appreciate this from the type of bond elemental nitrogen has. Oxides of nitrogen, which eventually are converted to nitrates become water soluble, making them available for plant usage. The plants, when eaten by animals pass the nitrogen to them. Synthesis of biologically important molecules such as amino acids and nucleic acids require nitrogen. Put students into two groups to

Competencies	*Contents*	*Suggested Activities*
		debate why fertilizers like urea and diammonium phosphate are extensively used. What role do the nitrogen-fixing bacteria in legumes play in the nitrogen cycle? Students should know that there is no gaseous form of pure or in a combined form of phosphorus. It is found only in water and land. Students should appreciate how overuse of phosphorus in the form of fertilizers and detergents has affected the environment negatively.
• define metallurgy, • explain the major steps in metallurgical processes,	**5.2 Some Elements in Industry** *(12 periods)* **5.2.1 Metallurgy** • Pretreating the ore • Converting the mineral to the element • Refining and alloying **5.2.2 Extraction, properties and uses of some selected metals**	Students should be able to see how metallurgy, the science and technology of metals, is used in the commercial production of metals. The three principal steps in metallurgy are concentrating and pretreating the ore, converting the mineral to the element by reduction and refining and alloying. Discuss each step why and how it is done and ask students to support their reasoning with chemical principles they already know.
• describe the manufacture of sodium by the Down's cell, • explain the chemical properties of sodium • describe the uses of sodium,	• **Sodium** - Extraction - Chemical properties - Uses	For all elements discussed in this section, chemical principles have to be highlighted why certain extraction methods but not others have been used. Ask first the students to identify the methods of extraction, the chemical they have and their uses for all six elements being considered here.
• describe the manufacture of calcium • explain the chemical properties of calcium, • describe the uses of calcium,	• **Calcium** - Extraction - Chemical properties - Uses	Students should know that among the alternative reduction methods, extraction of sodium and calcium uses electrochemical method of reducing fused NaCl and fused $CaCl_2$ respectively, but using coke (elemental C) does not work, why? Calcium can also be prepared by reduction of CaO by aluminum in a
• describe the manufacture of tin, • explain the chemical properties of tin, • describe the uses of tin, • describe the manufacture	• **Tin** - Extraction - Chemical properties	

Competencies	***Contents***	***Suggested Activities***
of lead, • explain the chemical properties of lead, • describe the uses of lead, • describe the manufacture of zinc, • explain the chemical properties of zinc, • describe the uses of zinc, • describe the manufacture of chromium, • explain the chemical properties of chromium, • describe the uses of chromium.	- Uses • **Lead** - Extraction - Chemical properties - Uses • **Zinc** - Extraction - Chemical properties - Uses • **Chromium** - Extraction - Chemical properties - Uses	vacuum, where the calcium produced distills off. Students should know the chemical properties and common uses of sodium and calcium. Students should know that tin, lead, zinc and chromium are extracted from their ores by using a reducing agent coke (elemental carbon) in furnace, unlike sodium and calcium which can't. They should also know that lead and zinc are produced in two steps. They are: • First, roasting: oxidation of sulfides to oxides. Because it is easier to reduce oxides. • Next, reduction: reduction of oxides to their elements using a reducing agent coke(elemental carbon) Students should know the chemical properties and common uses of Tin, lead, zinc and chromium.
• Describe production of silicon	**5.2.3. Silicon** • Production	Students should appreciate how

Competencies	*Contents*	*Suggested Activities*
• Explain the chemical properties of silicon • Describe uses of silicon	• Chemical properties • Uses	silicon is prepared by reducing quartz sand, SiO_2, with coke, C, in electric furnace at 3000 ^{0}C . The teacher should emphasize the semi-metallic nature of silicon and its modern uses based on semi-conductivity.
• Explain the steps in Haber process, the industrial production of ammonia and its uses • Explain the steps in Ostwald's process, the industrial production of nitric acid and its uses • Explain the steps in Contact process, the industrial production of sulphuric acid and its uses • Explain the steps in the industrial production of Diammonium Phosphate and its uses	**5.2.4. Some important compounds of selected nonmetals** • Ammonia • Nitric acid • Sulphuric acid • Diammonium Phosphate (DAP)	One could always select different molecules for discussion based on the criteria used for selection. But these ones are among the many considered to be important in today's world. In each one of them go through the following points: • elemental composition, structure, polarity, properties and uses • use the methodical approach and work your way by using the elemental composition as an explanation for the structure and so on

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms mineral , ore and metallurgy; describe the forms of occurrences of metals; discuss the distribution and relative amounts of the elements in the earth's crust; define the term fixation; discuss the carbon cycle, the nitrogen cycle and the phosphorus cycle; explain the major steps in metallurgical processes; describe the manufacture of sodium ,calcium, tin, lead, zinc and chromium; explain the chemical properties of sodium, calcium, tin, lead, zinc and chromium; describe the uses of sodium, calcium, tin, lead, zinc and chromium; explain the steps in Haber process, the industrial production of ammonia, the Ostwald's process, the industrial production of nitric acid, the contact process, industrial production of sulphuric acid, the industrial production of Diammonium Phosphate, and the industrial production of Silicates ;

Students above minimum requirement level

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Students below minimum requirement level

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Unit 6: Polymers *(14 periods)*

Unit outcomes: Students will be able to:

- understand the process of addition & condensation polymerization;
- list a variety of synthetic polymers & natural polymers & explain their differences;
- classify synthetic polymers as addition or condensation polymers & identify their monomers;
- describe the monomers, properties & uses of plastics, rubbers, carbohydrates & proteins;
- analyze the risks & benefits of the development & application of synthetic polymers(e.g. plastics) & suggest possible methods of control;
- describe scientific enquiry skills along this unit: classifying, comparing & contrasting, communicating, asking questions, applying concept and making generalization.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • define the terms monomer and polymer; • classify polymers into synthetic and natural polymers; • give examples of synthetic and natural polymers;	**6. Polymers**	Students should know that macromolecules are made of smaller units(monomers). They could be made of identical monomers- homopolymers or different types of monomers- copolymers. Polymers could be: • Synthetic • Natural Ask students to list down examples of polymers from their surroundings and classify them as synthetic or natural polymers.
• explain polymerization; • mention the two types of polymerizations; • explain how addition polymerization takes place; • explain how condensation polymerization takes place; • give examples of addition polymers; • tell the monomers of each of the given addition polymers • explain the common uses of addition polymers;	**6.1 Polymerization** *(2 periods)* • Addition polymerization • Condensation polymerization	Students should be aware that polymerization could be effected either through: • Addition polymerization or • Condensation polymerization Give students reading assignment on the difference between addition and condensation polymerization to be presented to fellow students. Help them identify the functional groups involved in each case. This is a good opportunity to show students how applied organic chemistry is in every day life both synthetically and naturally.
• give examples of condensation polymers;	**6.2 Synthetic polymers** *(5 periods)*	Students should know that the number of synthetic polymers is increasing rapidly in response to

Competencies	*Contents*	*Suggested Activities*
• tell the monomers of each of the given condensation polymers; • explain the uses of condensation polymers: nylon, polyester, and bakelite; • describe thermoplastics and thermo set polymers • give examples of thermoplastics and thermo set polymers;	• Addition polymers - Polyethylene - Polypropylene - Polyvinyl chloride - Polymethylmetacrylate (Perspex) - Polystyrene - PolyTetra Flouro Ethylene (PTFE) or Teflon • Condensation polymers - Nylon - Polyester (Dacron) - Bakelite • Thermoplastics and thermo set polymers	the needs of mankind. Ask if students know what problems are arising with this development. Do they have opinions about what needs to be done to protect the environment? You could tell them at this point there are fields of study known as Environmental Science and even Environmental Chemistry. Students should understand problems arise when plastics are discarded into the environment. Assign students to collect synthetic polymers which they think are useful or with serious drawbacks. Could they identify the polymers, the corresponding monomers and whether they are addition polymers or condensation polymers? You should appreciate the effort they put but be ready to answer the questions you asked them; they may not have the answers. You may not know what they have collected. It is not always easy. You could bring properly identified samples.
• describe natural rubber; • tell the monomer of natural rubber; • explain vulcanization • explain the use of natural rubber; • give three examples of synthetic rubbers; • explain uses the examples of synthetic rubbers; • compare and contrast synthetic and natural rubber;	**6.3 Natural polymers** *(7 periods)* • Rubber - Natural Rubber - Synthetic Rubber	Students should be aware one good thing about synthetic polymers is that they can design and produce materials with properties of their desire. Students may want flexible but of considerable strength or hard and brittle polymers. Students may want to reuse old polymers as starting materials for new ones; in other wards to recycle substances. Ask the students what uses recycling has. What makes thermoplastic polymers recyclable whereas thermoses ones are not? Ask students if they can see similarity about intermolecular forces in solids covered in grade 11 and thermoplastic and thermo

Competencies	*Contents*	*Suggested Activities*
		set polymers. There is a popular saying, 'Nature had it first'. Long before mankind started making synthetic polymers there were natural polymers everywhere. Ask students to list down as many natural polymers as they could find around them or even within them. Can they see that the number of natural polymers is almost constant but every day new synthetic polymers are made.? Students should know the monomers for both synthetic and natural rubber polymers. Discuss the role of sulfur in making rubber harder (vulcanization). You may bring this up when you discuss proteins that use similar bonding with cysteine amino acid.
• define the term carbohydrate • describe mono saccharides; • draw the structure of monosaccharide, • give examples of mono saccharides • describe disaccharides; • draw the structure of disaccharides; • give examples of disaccharides; • describe poly saccharides; • draw structures of starch and cellulose; • explain the difference between starch, glycogen, and cellulose;	• Carbohydrates - Monosaccharide - Disaccharide - Polysaccharide	Carbohydrates are compounds that have C, H and O. You may tell your students that carbohydrates are poly hydroxyl aldehydes or polyhydroxy ketones. Most carbohydrates are homopolymers of simple sugars known as monosaccharides. The monosaccharides are the monomers for polysachharides which could contain thousands of monomers. Most polysaccharides are homopolymers. Warn your students to pay special attention to the monomer, glucose, as it has a bigger role than any other monosaccharide in biochemistry. Ask students to draw various structures of few monosaccharides including glucose. In water solution the cyclic conformations are dominant. Ask the students whether there is any similarity between starch,

Competencies	*Contents*	*Suggested Activities*
		cellulose and glycogen. Which ones can be used as energy sources by mammalian cells? Do they give the comparable amount of energy if combusted in a crucible? Do you see reasons for the differences, if any? Discuss this from bonding point of view. Ask students if they can identify synthetic fibers used as clothing that were developed on principles learned from carbohydrates? Are carbohydrates condensation or addition polymers?
• define the terms amino acid, peptide and protein; • describe the structure of amino acids; • explain how proteins are formed; list types of proteins;	• Proteins - Amino acid - Polypeptides	Students should know that proteins got their names from a similar Greek word meaning of primary importance. They are made by polymerization of twenty naturally existing amino acids. Students should know why they are called amino acids. Discuss some of their properties that are important for polymerization. Try to show similarity in the way amino acids are polymerized to form proteins with nylon. Why do amino acids and monomers of nylon have two functional groups? Ask students to see that proteins are condensation polymers. What small molecule is released as a byproduct here? Limit the discussion to introducing that they are polymers of nucleotides, which themselves consist of an organic base, the ribose (or deoxy) sugar and at least one phosphate group. Ask the students if they already know the two types of nucleic acids, RNA and DNA. They are called nucleic acids because they were isolated and identified as acids because of the phosphate group, from the nucleus of eukaryotic cells.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: define the terms monomer and polymer; classify polymers into synthetic and natural polymers and give examples for each of them; explain polymerization and mention the two types of polymerizations; explain how addition and condensation polymerizations take place; give examples of addition polymers and tell the monomers of each of them; explain the common uses of addition polymers; give examples of condensation polymers and tell the monomers of each of them; explain the uses of condensation polymers: nylon, polyester, and bakelite; describe thermoplastics and thermo set polymers and give examples for each of them; describe natural rubber and tell its monomer; explain vulcanization and the uses of natural rubber; give some examples of synthetic rubbers and explain their uses; compare and contrast synthetic and natural rubber; define the term carbohydrate; describe monosaccharide ,draw its structure and give examples; describe disaccharides, draw their structure and give examples; describe poly saccharides; draw structures of starch and cellulose; explain the difference between starch, glycogen, and cellulose; define the terms amino acid, peptide and protein; describe the structure of amino acids; explain how proteins are formed; list types of proteins.

Students above minimum requirement level

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Students below minimum requirement level

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