CHEMISTRY

TEACHER'S GUIDE

GRADE 11

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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 The traditional teaching-learning paradigm has been replaced by active, participatory and student-centered model.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 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:

i. the concepts they gained during discussion in each activity in all units.
ii. their observation and analysis after performing experiments in groups to the class.
iii. 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 teachers 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 11 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 the information on fundamental concepts of chemistry.

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 constructing a model of 1,2-dichloroethene.

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 (peer teaching)

This is an activity where students present a topic in front of their classmates. This can be done individually or as a group.

E. Demonstration

This is a method where the teacher shows the students how something is done.

Experiments

It usually involves a very specific and controlled method of procedures, where results are usually recorded. This method is applicable in performing laboratory experiments throughout units 2 – 6 at this level.

F. 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, classification major branches of chemistry.

G. Question and answer (inquiry)

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

H. Investigation

This method is usually based on real life. For example, the effect of hydrogen bonding on aquatic animals.

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.

J. Visual-based active learning

This method helps students learn using real object models, pictures, drawings and charts. For example, this method can help in teaching the valence bond method and the molecular orbital method etc.

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 while teaching the international units measurement.

L. Problem solving

Problems solving activities involve students finding solutions to problems. Problem solving can be done individually or in groups. The solution is not the focus. Instead, students are encouraged to explore different strategies and processes to find the solution. It creates students who are able to think for themselves or independent thinkers and look for solutions rather than become trapped in problems.

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.

UNIT 1 FUNDAMENTAL CONCEPTS OF CHEMISTRY

UNIT OVERVIEW Total Period allotted: 15 periods

This unit reviews basic mathematical techniques and measuring principles that are important in problem solving and laboratory work. In addition, the factor-label method of problem solving is introduced and used throughout the unit.

SI measurement units are used throughout the text with few exceptions. SI is used to globally facilitate communication among people all over the world.

Unit outcomes

After completing this unit, students will be able to:

- *understand the scope and major fields of chemistry;*
- *select and use appropriate SI units and prefixes;*
- *understand the causes of uncertainty, precision and accuracy in measurement;*
- *express the results of any calculation involving experimental data to the appropriate number of decimal places or significant figures;*
- *use scientific notations to avoid ambiguity and to make it easier to understand;*
- *know how to round off significant figures using rounding off procedures;*
- *understand how to use significant figure in calculations and dimensional analysis as an aid in problem solving;*
- *use scientific method in solving problems;*
- *demonstrate an understanding of experimental skills in chemistry;*
- *demonstrate a knowledge of laboratory procedures.*

Main Contents

1.1 The scope of chemistry

1.2 Measurements and units in chemistry

1.3 Chemistry as experimental science

Introduction

The students' text has examples, exercises, activities and experiments that are intended to provide supplementary stabilizations to their knowledge in chemistry.

Chemistry is a complicated and tough science. However, students in Ethiopia, for the most part, that reputation is undeserved. Although the mathematics and logic need to apply when they take classes in chemistry, can make studying chemistry a challenge for them if they are weak in those areas, anyone can understand the basics of how things work... and that's the study of chemistry. Chemistry is the science of matter - what matter is like, what makes up matter, and what changes occur in matter, what causes those changes. It is a key to all the other sciences and a part of every aspect of our world. In a nutshell, chemistry explains the world around us. Therefore, it is the duty of chemists in general, and chemistry teachers in particular, to raise the interest of students in chemistry.

Chemistry explains how food changes as we cook it, how it rots, how to preserve food, how our body uses the food we eat, and how ingredients interact to make food. It explains how cleaning works. We use chemistry to help decide what cleaner is best for dishes, laundry, ourselves, and our home. We use chemistry when we use bleaches and disinfectants and even ordinary soap and water. How do they work? That's chemistry! We need to understand basic chemistry so we can understand how vitamins, supplements, and drugs can help or harm. The importance of chemistry also lies in developing and testing new medical treatments and medicines. Chemistry is at the heart of environmental issues. What makes one chemical a nutrient and another chemical a pollutant? How can we clean up the environment? What processes can produce the things we need without harming the environment?

All of us are chemists. We use chemicals every day and perform chemical reactions without thinking much about them. Chemistry is important because everything we do is chemistry! Even our body is made of chemicals. Chemical reactions occur when we breathe, eat, or just sit reading books, journals, tabloids, newsletters and the like. All matter is made of chemicals, so the importance of chemistry is that it's the study of everything.

Start-up Activity

The purpose of this start-up activity is to motivate students for active participation during discussions on the definition of chemistry. Therefore, form groups of students and engage them in discussing questions raised and them in presenting the consensus reached by the groups to the rest of the class.

1.1 THE SCOPE OF CHEMISTRY

Periods allotted: 2 periods

Competencies

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

- *define chemistry;*
- *distinguish the major fields of chemistry;*
- *distinguish the subdivision of branches of chemistry.*

Forward Planning

Generally teachers should update themselves because experts tend to forget the subject matter if they stop reading for a period of time. To play the role of a model teacher, discharge the following duties and responsibilities:

- Note what is contained in the book related to the topic you are intending to teach.
- Thoroughly read and understand the content chosen for the daily lesson.
- If it appears to you that the textbook material is not enough for the daily lesson you are planning to teach, supplement it with references.

Teaching Aids

You may use flipcharts or flash cards to highlight the key words used in this section.

Subject Matter Presentation

Definition of Chemistry

Our students seem to lack interactive competence. Therefore, they should learn and develop this important skill. To help them to do so, we advise you to devise an interactive teaching-learning method in which students express ideas and learn ways of expressing valid arguments. For example, you can use group-discussion and role-playing methods to serve this function.

This section is mainly concerned with the definition of the role of chemistry in development, the definition of terms, physical and chemical changes, and major fields of chemistry and with explaining possible links that exist between chemistry and related sciences. The teacher has to familiarize himself with the various definitions from different angles and be able to cite references and provide practical examples.

To introduce this topic, you could motivate your students by stating that the knowledge of chemical substances is essential to the study of chemistry. The principle of chemistry that we consider in the textbook has many useful applications. In some cases, exploring the practical side of chemistry has stimulated the discovery of new principles. In chemistry, theory and application are intertwined.

Then you could continue the lesson by involving students in discussions, using *Activity 1.1.* The purpose of this activity is to illustrate the role chemistry plays in industry and in society as well. Whether they are aware of it or not, people are always aided by chemistry in their day-to-day activities. The materials people use everywhere in the universe are directly or indirectly the results of chemistry. After the discussion, explain your own view about this activity by mentioning for example:

Toothbrush

Usually made of a plastic handle with nylon or polyester bristles attached to form a brush. Other options are available, however, using natural or biodegradable products which are more environmentally friendly.

Toothpaste

Modern toothpaste has many things to do. It must have abrasives to scour off bacterial films. It must have fluorides to harden the teeth against decay. It must have a strong enough flavor to hide the bad tastes of decaying bits of previous meals, and the awful taste of some of the other ingredients, in the paste such as detergents and phosphates.

Toothpaste must have thickeners to stay on the toothbrush, and must squeeze out of the tube. It must have detergents to remove fatty films, and needs water softeners to make the detergents work better. It also needs sweeteners. Preferably those are non-nutritive, so bacteria are not encouraged.

Fluorides

This compound prevents tooth decay, and makes teeth stronger.

Tetra sodium Pyrophosphate anti-tartar agents: Pyrophosphates stop formation of tartar by removing tartar elements from the saliva.
Sodium Bicarbonate (baking soda): This is an abrasive, and provides a mild tooth-whitening action.

Zinc citrate or zinc chloride: is used as an antibacterial agent to control dental plaque. The combination of zinc chloride and triclosan is effective against bacteria and gum disease.
Potassium nitrate, strontium chloride, potassium citrate (desensitizing Agents): Some people experience acute pain while eating sweet or frozen desserts, such as ice cream.

This happens due to tooth-sensitivity problems. Toothpastes containing desensitizing agents help to relieve discomfort and pain that accompanies sensitive teeth. Hydrogen Peroxide (whitening or abrasive agent): It performs the function of polishing and cleaning the teeth. It is used in toothpastes to remove stains on the surface of teeth.

Soap

Soap is made from long-chain carboxylic acids derived from animal fat or plant oils.

Soaps and detergents are used to remove all kinds of dirt, including grease, fats, oils, sweat and blood stains. Grease and water don't mix, so the cleansing action of detergents and soaps requires special chemistry.

The principle of "like-dissolves-like" suggests that soap or detergent molecules must include one portion that is grease-soluble and another that is water-soluble. Although the chemistries of soaps and detergents are somewhat different, the basic mechanism remains the same. The fat-soluble portion of soap or detergent combines with greasy dirt particles. The water-soluble part surrounds or encases each particle in a sphere. This sphere dissolves in the wash water and is disposed of.

Petrol gives a great deal more heat energy than does wood. Also, it helps natural conservation.

Jeweler

From the jeweler’s point of view, gold and silver are the two most-used metals in jewelry. The two attributes of gold which make it an excellent metal in which to work are its malleability and ductility. Indeed, gold is the most malleable and ductile of all metals.

Gold is also preferred for making jewelry because it does not react with atmospheric moisture and does not rust. Furthermore gold has a stable market value that is almost always rising. And its cost per weight is high, making the jewelry expensive and thus lucrative to the jeweler. It is a precious metal that is highly valued by most people over the world, and has been even since prehistoric times! Its bright yellow color and the way that it reflects light make it very attractive. It also does not rust, or corrode, is easy to clean, and it is very durable. Therefore, many or most people feel that gold is undoubtedly the best metal for jewellery.

Cooking

Cooking makes the food more attractive and appetizing, makes the food easy to digest, and serves to kill the germs and parasites that might have found places in the food. And also if you don’t cook meat... if it is raw and you just eat it... you may get very sick!

Encourage some of the students to propose definitions of chemistry in their own words. In listening to students' response, focus on the key words used in the students' attempt. In your discussion that follows, harmonize the background knowledge of your students with the refined definition in their textbook. Discuss terms like composition, structure, transformation, and properties of matter which may be used in the definition of chemistry. Conduct a spotted discussion with the students to create a common understanding of how matter and our environment are regarded in chemistry.

Major Fields of Chemistry

After the students are familiar with the definition of chemistry, you could proceed to discuss the major fields of chemistry. Here again, organize students in groups and guide them to perform *Activity 1.2* which is thought to assist the mastery of major fields of chemistry. This activity can be given as homework to be done using available reference materials. On your part you need to prepare yourself for supporting students, as they may require your assistance in some of the ambiguities they may face in their attempt to do the activity.

1. Physical chemistry, for example, can be subdivided into thermodynamics, quantum chemistry, kinetics and electrochemistry. Guide students to perform in the same way.
2. Chemistry in daily life follows as we bathe, cook and wash. When we bathe, we use soap made out of animal fats or vegetable oil and some chemicals, applying the knowledge of chemistry. Chemistry can be found in practically every aspect of modern life. Chemists invented cheap dyes for clothes, polymers in plastics, the art of glass making, metal extraction processes for aluminum and iron, batteries, drugs such as paracetamol and even alternatives for sugar.

It is amazingly important - there is an aspect of chemistry in everything - from the plastic of your keyboard to the LCD in TVs to the rubber tires of cars, from the molten centre of the earth to the ozone layer.

Additional Notes

To illustrate the inter linkage of theory and application in chemistry, you may consider some issues concerning, for example, chlorine, one of the most familiar chemical elements.

Today, the chemical industry turns out more than 10,000 chlorine compounds ranging from polyvinyl chloride (pvc) for pipes to compounds used as bleaches, disinfectants, solvents, pesticides, and drugs. Water-borne diseases are eliminated when water supplies are treated with chlorine to kill pathogenic organisms.

Trace amounts of toxic compounds called **dioxins** are produced by the combustion of chlorine-containing plastics and by the use of chlorine as a bleaching agent in the pulp and paper industry. Another example of chlorine compounds that cause environmental concern are the chlorofluorocarbons (CFCs), used in refrigerators and air conditioners and in making foamed plastics.

Chemical knowledge was used to make the chlorine products that have provided so many benefits, chemical knowledge played a large role in revealing the negative side of these products, and even more chemical knowledge will be required in the search for suitable substitutes.

Assessment

Assessment must be continuous. Do not limit your evaluation to mid and final exams. Ask your students (1) to define chemistry, making sure that the definitions they give reflect real understanding of the meaning but not mere imitations. (2) to explain possible relationships that may exist between chemistry and other natural sciences. An equally important point involves asking them to list and explain physical and chemical properties of substances, including intensive and extensive properties, with examples.

Exercise 1.1 will also help you for this purpose. You may help some of the students that have difficulty in answering the questions. You may use notes to record the performance of your students, and to identify the understanding level of your students.

Additional Questions

* 1. Write the definition of chemistry?

2. What is the scope of chemistry?

Answers to Additional Questions

1. Chemistry is the science of the composition, structure, properties, and the reaction and the elements of a complex entity and their dynamic interrelation. Or Chemistry is a science that describes the physical properties, composition, structure and reactions of matter.

2. The scope of chemistry is huge. The days are long past when one person could hope to have a detailed knowledge of all areas of chemistry. Those pursuing their interests into specific areas of chemistry communicate with others who share the same interests. Over time, a group of chemists with specialized research interests become the founding members of an area of specialization. The areas of specialization that emerged early in the history of chemistry, such as organic, inorganic, physical, analytical, and industrial chemistry, along with biochemistry, remain of greatest general interest.

Based on some scholars' views, Biochemistry takes the highest percentage, inorganic chemistry is the second, and then applied chemistry, chemical engineering, organic chemistry, physical chemistry and analytical chemistry follow.

Answers to Exercise 1.1

1. Physical changes are easily reversible. They are not accompanied by any marked external effects. Usually no change of mass occurs and they produce no new kind of matter. It is not accompanied by great heat change (except latent heat effects) accompanying changes of state. Chemical changes always produce a new kind of matter. They are usually accompanied by considerable heat change, and they produce individual substances whose masses are different from those of the original individual substances.
2. Intensive properties include specific gravity, freezing point, smell (odor), solubility, conductivity. Extensive properties: Thickness, area, weight of a substance.
3. Students may refer to textbook.

Answers to Exercise 1.2

1. Yes, the three state of matter are inter-convertible. These changes are physical changes.
2. The students are expected to refer to other chemistry textbooks as well as any source they find in their local areas and present their findings to the class for discussion.

1.2 MEASUREMENTS AND UNITS IN CHEMISTRY

Periods allotted: 8 periods

Competencies

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

- *list and describe the seven SI units and their prefixes;*
- *write the names and symbols of derived SI units;*
- *use the factor label method for solving problems and making conversion of SI units;*
- *describe uncertainty of measurements;*

- *identify the digits that are certain and the ones that are uncertain, given a number representing measurement;*
- *define precision and accuracy;*
- *estimate the precision possible for any instrument you use in the laboratory;*
- *explain system and random errors;*
- *analyze a given data in terms of precision and accuracy;*
- *define decimal places;*
- *determine the number of decimal places in a calculated result;*
- *define significant figures ;*
- *determine the number of significant figures in a calculated result; and*
- *use the scientific notation in writing very large or very small numbers.*

Forward Planning

Teaching SI may require that you first familiarize yourself with the system. You can use supplementary materials to assist the teaching-learning process.

This section mainly deals with how data can be collected by chemists, and the data collected is used in computations. You can prepare countable items, measuring instruments such as balances, volumetric glassware, thermometers, rulers, clocks, etc,.

You should familiarize yourself with a wide variety of units of measured quantities. For example, volume could be measured and expressed in units such as liter, cubic centimeter, gallon, quart, cubic yard, etc. Also, devise a technique to investigate the students knowledge of SI units.

Teaching Aids

Read activities suggested under this topic thoroughly and prepare the materials required for yourself and let the students be ready ahead of time. Prepare different objects with different mass, thermometers, meter sticks, graduated cylinders and also watches and balances of the same sensitivity.

Subject Matter Presentation

SI Units (The International System of Unit)

You can apply teaching methodologies such as discussion, experiment and independent work. In introducing this section, you may start discussion by emphasizing the importance of quantitative data in chemistry. Let your students notice that measured quantity without a unit attached to it is meaningless. For example, the distance between Addis Ababa and Awassa is 275. This statement is meaningless unless units like km, mile, meter, etc are attached to it. Measurement, therefore, must be expressed both in numbers and units. Help students, to have a combined idea of English unit of measurements and SI units, of course giving emphasis to the SI units. This approach is preferable because students in their future careers are likely to encounter these non-SI nits. So do not hesitate to introduce the common English units. See the following examples of SI and English units:

Units of Mass	**Units of Volume**	**Units of length**
1 kg = 2.205 lb	1L = 1000 cm^3 = 1.057 qt	1 m = 39.37 in
1 lb = 453.6 g	1 m^3 = 1000 L	1 in = 2.54 cm
28.35 g = 1 OZ	1 gal = 3. 785 L	30.48 cm = 1 ft
1 metric ton = 1000 kg = 1.102 tons		1.609 km = 1 mi

Introduce that all numbers obtained as a result of counting carry the name of the item. For example, 18 students in a class, 602 atoms of hydrogen, 5 oranges in a basket etc. Numbers that are obtained as a result of measurements taken should be associated with a unit. For example, 12 kg of sugar, a 21cm stick, 56 seconds to leave etc.

For better illustration, classify your students into groups and ask each group to perform one of the following measurements. Tell the students to perform each measurement very carefully. Let each student in a group measure and report accordingly:-

1. the mass of any given object, using any available balance, in grams;
2. the room temperature (in ^{o}C) of their working room;
3. the length of an armchair in their classroom, using a ruler or meter stick;
4. 16 mL of water, using a graduated cylinder;
5. the time required to count numbers from 1 up to 100 (in seconds).

For the same group, give *Activity 1.3* and encourage the students to participate actively and report to the class. This activity is intended to familiarize students with SI units and measuring devices. After group discussion and presentations, briefly explain the way you view the activities, which may or may not be the same as the students' presentation. For example, a meter stick is used for measuring length, a beam balance for measuring mass, a measuring cylinder for measuring the volume of HCl, a stopwatch to measure the time taken for a reaction, and a thermometer to record the temperature. The units used to express these physical quantities might be centimeter, gram, milliliter, minutes and degree centigrade, respectively.

Mass is the amount of matter in a body. Length is the measurement of matter from one end to the other. Heat is an energy transfer into or out of a system that is caused by a temperature difference between two points. You can say that temperature is “hotness”, though it is not very precise. From common experience, it is known that if two objects at different temperature are brought together, heat flows from the warmer to the colder object.

Activity 1.4 is assumed to give a clear understanding of the difference between mass and weight. Mass measures how much matter presents in an object. In other words, mass is a measure of the amount of matter an object possesses. Weight is the heaviness of matter, especially as measured against gravitational force. In contrast to weight, mass is not affected by gravity.

Derived SI units

Let your students notice the difference between basic SI units and derived SI units. The in-text questions are helpful in preparing students to participate actively in the lesson. Several examples are given in the text to illustrate the mathematical relations described in this section. Let the students study independently the examples given. You will support them whenever there is ambiguity in their understanding.

Emphasize dimensional analysis. You may prefer to use the phrases unit analysis or conversion factor for dimensional analysis. Unit analysis is not mentioned in students’ textbook. Investigate the knowledge of students, emphasizing:

- ✓ the importance of cancellation of unit
- ✓ the relationship between the conversion factor and its inverse
- ✓ the technique of keeping track of the status of a series of conversions
- ✓ the requirement, at times, to convert units in both the numerator and the denominator of an expression.

Stress some of the common prefixes used in the SI unit of measurement followed by *Activity 1.5* which enables students to have the idea of how reagents are taken in laboratories for experiment and compare with those of industrial scales. Experiments in the laboratory should be performed with small amounts of reagents for many different reasons like: reducing hazards which may be caused by violent reactions when concentrations increase, economy, handling, ease of performance, time factor, etc.

Uncertainty in Measurement

You may use *Activity 1.6* in giving the lesson on uncertainty in measurements. Feel sure, you will come across different reports on the results of their measurements. This will help you to discuss with the students the impossibility of obtaining completely perfect measurements. I.e., it is impossible to obtain the exact value of any measurement. Indicate a margin of error when you come to the discussion on significant figures.

Precision and Accuracy

Your discussion on precision and accuracy can be facilitated by students' work on *Activity 1.7*. Help students to pay attention and discuss on the fact that precision and accuracy are two important terms that describe reliability of measurements. Help them to have a clear idea of the distinction of these two terms – precision and accuracy. Use Example 1.2 for stabilization of the concept. The following suggested answers may be used for *Activity 1.7*.

a.

Student	***Average***
A	(8.72 g + 8.74 g + 8.70 g)/3 = 8.72 g
B	(8.50 g + 8.77 g + 8.83 g)/3 = 8.70 g
C	(8.50 g + 8.48 g + 8.51 g)/3 = 8.50 g
D	(8.41 g + 8.72 g + 8.55 g)/3 = 8.56 g

It can be seen that the average mass of student A is exactly equal to the true value. Therefore, student A's is the most accurate.

b. The measurements of student C are close to each other, and his measurements are the most precise. It is not the most accurate.

c. The measurements of student C give the least accurate data (whose average is 8.50 g) – far from the true value. No, it is the most precise.

The precision of measurements can be evaluated in terms of average deviation and relative average deviation from the average values as:

$$\text{Average deviation} = \sum \frac{/\text{indivdual value} - \text{average value}/}{\text{number of measurements}}$$

Note: The lower values for average deviation and relative average deviation indicate more precision i.e, less uncertainty.

Therefore, for student A: /8.72 – 8.72/ + /8.74 – 8.72/ + /8.70 – 8.72/ = 0.04

Average deviation = 0.04/3 = 0.013

Relative average deviation = 0.013/8.72 x 100 = 0.15

Similarly, for students B, C, and D, the average deviations and relative average deviations can be calculated in the same way as student A.

Student	Average deviation	Relative average deviation
B	0.13	1.5
C	0.01	0.12
D	0.11	1.25

Decimal Places

Though the idea of decimal is not new for these grade level students, emphasis must be given to the rules in expressing decimal numbers and rounding off. Give special attention to rule number three in the students' textbook.

Significant Figures

It is important to develop the use of significant digits. Using a calculator or computer requires that students apply some rules in making calculations in order to obtain answers with reasonable significant figures. There can be some disadvantages to being too strict in assessing students' answers. It is important that they understand a number of mathematical relationships and can perform operations based on them. Being too critical of the number of digits in the answer can be frustrating for students. As a result of these limitations, help the students to comply with the rules on how to express significant figures because a calculated quantity is no more precise than the data used in the calculation.

Organize students in groups and let them discuss significant figures. At the end of their discussion ask them to describe how uncertainty and precision, as well as significant figures and precision, are related. You can extend the discussion based on some exercises which are relevant to the topic discussed.

Assessment

Don't forget to assess students' active involvement in each of the activities. Each of the exercises in this section should be properly used to assess the students' mastery of each concept in the sections. Your student performance record list is helpful in all your assessments. Don't forget to appreciate students working above minimum requirement level. You are required to assist those working below the minimum requirement level.

Additional Questions

* 1. Convert 1.2 cm to µm.

2. How many significant figures are there in:

a. 204 kg　　b. 0.004 g

c. 0.0802 cm　　d. 90560 N

3. Convert 86 ng to mg.

* 4. Use scientific notations to express:

a. 204 kg b. 0.004 g

c. 292689 km d. 0.0802 cm

* 5. Express the result of adding 856.6g and 23.86g to the correct number of significant figures.

6. Express the result of adding 0.203 cm and 22.7 mm to the correct number of significant figures.

* 7. Round off each of the following to three significant figures.

a. 20605 g b. 0.04090 mL

c. 9756 m d. 1.675 s

8. What is the total volume produced when 1.037 of water is mixed with 5.440 mL of water?

9. The answer to the expression 8.735 g + 2.3 g + 93.683 g, to the correct number of significant figures, is

a. 104.7 b. 104.718 g

c. 1.05×10^2g d. 104.6 g

* 10. For 2.05×10^2Å, the equivalent expression in meters is

a. 2.05×10^{12} m b. 2.05×10^{-10} m

c. 2.05×10^{-12} m d. 2.05×10^{-8} m

* 11. The difference between an accepted value and an experimental value is called a(n):

a. Error b. Percent Error

c. Measured Value d. Precise Measurement

Answers to Additional Questions

1. Recall that 1cm = 10^{-2} m and 1μm = 10^{-6} m. Consequently

$$\frac{1\text{cm}}{1\mu\text{m}} = \frac{10^{-2}\,\cancel{m}}{10^{-6}\,\cancel{m}} = 10^4 \text{ or } 1 \text{ cm} = 10^4\ \mu\text{m}$$

2. a. Three: zeros between non-zero digits are significant.

   b. One: zeros preceding the first non-zero digit are not significant.

c. Three: for the reasons explained in (a) and (b).

d. Ambiguous - four or five: The final zero may or may not be significant. The first four digits, including the first zero, are significant.

3. The prefix nano (n) is 10^{-9}, so $1ng = 10^{-9}g$. And you know that $1mg = 10^{-3}g$.

So, $\frac{1ng}{1mg} = \frac{10^{-9}\cancel{g}}{10^{-3}\cancel{g}} = 10^{-6}$ or $1ng = 10^{-6}mg$

$$86ng = 86\cancel{ng} \times \frac{10^{-6}\,mg}{1\cancel{ng}} = 8.6 \times 10^{-5}\ mg$$

4. a. 2.04×10^{2} kg (just count how many places to the left you have to shift the decimal to give a number between 1 and 10, and this gives the exponent)

b. 4×10^{-3} g (just count how many places to the right you have to shift the decimal to give a number between 1 and 10, and this gives the negative exponent)

c. 2.92689×10^{5} km

d. 8.02×10^{-2} cm

5. 856.6 (significant to tenths of a gram)

$$\begin{array}{r} 856.6 \\ \underline{+\ 23.85} \\ 880.46 \end{array}$$

The result is significant only to tenths of a gram, so you round it off to 880.5 g.

6. You can't add apples to oranges-one of these measurements has to be converted to the units of the other. Since $1cm = 10^{-2}m$ and $1mm = 10^{-3}m$, you get

$\frac{1\ mm}{1\ cm} = \frac{10^{-3}\cancel{m}}{10^{-2}\cancel{m}} = 10^{-1}$ or $1mm = 10^{-1}cm$ and you can add the two quantities:

0.203 cm (significant to a thousandth of a centimeter)

\+ <u>2.27 cm</u> (significant to a hundredth of a centimeter)

$2.47\cancel{3}$cm

The total can be significant only to a hundredth of a centimeter, so you must report it as 2.47 cm (or 2.47 cm × 10mm/cm = 24.7 mm).

7. a. 2.06×10^{4} g b. 4.09×10^{-2} mL

c. 9.76×10^{3} m d. 1.68 sec

8. 1.042 L (or 1042mL)

9. a. 104.7 b. 104.718 g
   c. 1.05×10^2g d. 104.6 g

10. D

11. A

Answers to Exercise 1.3

1. 298.15
2. 150.8 ^{0}F

Answers to Exercise 1.4

1. 0.634 L
2. 7.47 g cm^{-3}
3. 10.5 g mL^{-1}
4. The density of mercury is greater than that of water, and oil has less density as compared to water.

Answers to Exercise 1.5

1. 5.22×10^4 g
2. 0.8 μm
3. 7.5×10^{-5}
4. 6.75×10^9
5. a. 3.22 μs b. 9056 mm c. 1.07 kg
6. 633 g

Answers to Exercise 1.6

a. 34.41 because it is the average of the four measurements.

b. They are neither precise nor accurate.

Answers to Exercise 1.7

c, d, e, and f

Answers to Exercise 1.8

1. 1.3×10^{-9}
2. a. 5.00×10^3 b. 3.986×10^3

Answers to Exercise 1.9

1. a. 415 g b. 6.3 c. 100.5
   d. 150 e. 45 f. 13
2. 1.54×10^{-7} mm
3. 9.92×10^{2}

1.3 CHEMISTRYAS EXPERIMENTAL SCIENCE

Periods allotted: 5 periods

Competencies

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

- *define scientific method;*
- *describe the major steps of the scientific method;*
- *use scientific methods in solving problems;*
- *demonstrate some experimental skills in chemistry*
- *describe the procedures of writing laboratory report.*

Forward Planning

This section deals with chemistry as an experimental science. A complete understanding of the scientific method, experimental skills in chemistry and writing a laboratory report is extremely important prior to presenting the topic to students. Devise, plan and carry out simple activities stated in the textbook.

Teaching Aids

Diagrams are required to show some schemes and equipment. Students need to visit the laboratory and observe the equipments available there. Read activities suggested under this topic thoroughly and prepare the materials required for yourself and let the students be ready ahead of time.

Subject Matter Presentation

The Scientific Method

In giving the lesson on scientific method, a diagram of Figure 1.5 of the students' text prepared as a flip chart is helpful. The lesson can be well established if the students are actively involved in *Activity 1.8.* You may use the following hints to facilitate the discussion.

Hold equal-sized plastic bags stuffed with different items of your interest and ask questions like.

- ✓ What things might you want to know about this bag?
- ✓ How many objects are in the bag?
- ✓ How much does the bag weigh?

Students will guess the number, the kinds of items in the bag, the weight, etc. Using the framework of the students guesses, you introduce scientific terminologies. For instance, during the discussion of answers to the question of how many objects (items) are in the plastic bag, the numbers put forth are hypotheses. At this point, you define the scientific term. In addition, write hypothesis on the black board, list the numbers volunteered by students underneath the word. After asking many if not all of the groups to respond with a hypothesis about how many objects are in the bag, you ask "how do you determine which hypothesis, if any, is correct?" Usually, students will ask you to open the bag. You then introduce the concept of **data collection** to determine if one's hypothesis is correct, etc.

Activity 1.9 can be performed using the role-playing method of presentation. In its more restricted modern sense, science is a system of acquiring knowledge based on scientific method, and knowledge gained through research. Science in this modern sense is a systematic way of gathering knowledge about the world and organizing and condensing that knowledge into testable laws and theories. Scientific knowledge is durable, Science demands evidence, science is a blend of logic and imagination, science explains and predicts. Science is not authoritarian. The sole criterion of scientific truth is experiment.

Some of the topics can be given to students in groups as reading assignment that are then to be presented in class by group leaders.

Some Experimental Skills in Chemistry

When you treat section 1.3.2, let students discuss the issues in *Activity 1.10.* Allow some students from their group to present their opinions. You can build your next explanation on the framework of students' presentations.

Facilitate and guide students to:

a. recognize the results of reactions of compounds

b. recognize the chemical tests for simple gases like: hydrogen, oxygen, carbon dioxide, ammonia, chlorine, nitrogen dioxide and sulfur dioxide

c. carry out simple chemical reaction experiments of your interest

d. apply the chemistry of the elements and compounds which includes the chemistry of Groups 1, 2, and 7

Observe and interpret details of the chemistry of the elements and compounds. Reference to standard texts is assumed for the interpretation of the results from observation exercises. Some of the above-mentioned experimental skills can be stabilized through *Activity 1.10*. Care must be taken here to avoid possible dangers and discomfort.

Activity 1.10 is assumed to help students in developing observation skills and exercising precautions during experiments. Nitric acid is known by its most pronounced oxidizing capacity and oxidizes most metals, including copper. The reactant copper disappears and produced in the form of copper (II) nitrate salt. Now the copper atoms are in the copper (II) nitrate salt. The other product accompanying this reaction is a toxic gas, NO_2. This change, as you know, is a chemical change.

$$Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2\,(aq) + NO_2(g) + H_2O(l)$$

Writing a Laboratory Report

Inform the students that there is a formal and conventional way of writing laboratory reports. Give them some time to read their textbook and then help them to write short and precise reports about their observations in *Activity 1.9.*

Assessment

Don't forget to assess students' active involvement in each of the activities. Each of the exercises in this section should be properly used to assess the students' mastery of each concept in the sections. You student performance record list is helpful in all your assessments. Appreciate students working above minimum requirement level. Don't forget to assist those working below the minimum requirement level.

Additional Questions

1. Does science stand on its own without any support? Is it the only way to know anything?
2. How is a hypothesis tested?
3. What must one be able to do with a theory?
4. Can a theory be proved?
5. What is experimental science?
6. Why chemistry is considered a science for everybody?
7. Who is known as the father of experimental science?

Answers to additional Questions

1. The answer is an unequivocal *no*. Since science must be prevented from being weakened by disciplines *other* than science, for science cannot pull itself up by its own without help from other disciplines. Science is like the second story of a house; it cannot stand without the first story and the foundation underneath.
2. You make an experimental prediction that would be true if the hypothesis is true. You do an experiment, and then conclude.
3. A scientific theory is a unified set of principles, knowledge, and methods for explaining the behavior of some specified range of empirical phenomena. By using scientific theories, one must be able to observe, sense, understand and experience the world around him and to attempt to explain how the natural world works. A scientific theory must have some logical consequences we can test against empirical facts by making predictions based on the theory.
4. The fact that a theory passed an empirical test does not prove the theory. The greater the number of severe tests a theory has passed, the greater its degree of confirmation and the more reasonable it is to accept it. However, to confirm is not the same as to prove logically or mathematically. No scientific theory can be proved with absolute certainty.
5. Science that uses what is known to try to prove ideas and concepts that are as yet untested.
6. Because chemistry is for everybody.
7. Galileo Galilee was and is sometimes referred to as "the father of experimental science." Galileo didn't take much on faith, rather, he tested his ideas through experiments.

Answers to Review Questions and Problems

PART I

1.	D	2.	B	3.	B	4.	D	5.	B
6.	C	7.	A	8.	C	9.	D	10.	C

PARTII

1. -40^{o}F
2. 45.7
3. 480 sec
4. a. 3.76 km b. 6.34 μg c. 1.09 ng

5. a. four b. four c. three
   d. five e. two f. four
6. a. 9.0×10^{-4} cm b. 2.000×10 sec
   c. 9.000×10^{3} sec d. 2.800×10^{3} m
7. a. 100.5m b. 436 g c. 111 ml
   d. 40 e. 3.5×10^{-4}
8. 3.12 g/ml
9. 5.24 g/cm^3
10. 11.3 g/cm^3
11. 11.3
12. 1 kg = 1000 g
    1 L = 1000cm^3

$$\frac{1000\text{ g}}{1000\text{ cm}^3} = \text{g / cm}^3 = \text{g cm}^3$$

13. Physical
14. The shortest is 0.0128 km, and the longest 17931 mm.
15. Meter

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UNIT 2

ATOMIC STRUCTURE AND THE PERIODIC LAW

UNIT OVERVIEW **Total Periods allotted: 28 periods**

In this unit, the concept of atomic structure and periodic law is discussed. Emphasis is given to the modern explanation of the atom.

Section 2.1 deals with historical development of the atomic nature of substances; Section 2.2 with Dalton's atomic theory and the modern atomic theory; Section 2.3 with early experiments to characterize the atom; Section 2.4 with the makeup of the nucleus; Section 2.5 with electromagnetic radiation (EMR) and atomic spectra; Section 2.6 with the quantum mechanical model of the atom; Section 2.7 with electronic configurations and orbital diagrams; and, finally, Section 2.8 deals with electronic configurations and the periodic table of the elements.

In all sections, many examples and exercises are included for the purpose of giving the students more chance to practice. Please advise students to take time and study them critically. This unit requires a minimum of 28 periods to complete it successfully. We advise you to do the best you can do while teaching this unit.

Unit Outcomes

After completing this unit, students will be able to:

- *understand the historical development of atomic structure;*
- *explain the experimental observations and inferences made by some famous scientists to characterize the atom;*
- *list and describe the subatomic particles;*
- *explain the terms atomic mass and isotope;*
- *understand electromagnetic radiation, atomic spectra and Bohr models of the atom;*
- *do calculations involving atomic structure;*
- *describe the quantum mechanical model of the atom and the related postulates and principles;*
- *demonstrate an understanding of periodic law and how electronic configurations of atoms are related to the orbital diagrams and be able to explain periodic trends;*

- *describe scientific inquiry skills used for this unit: inferring, predicting, classifying, comparing, and contrasting, communicating, asking questions and making generalizations.*

Main Contents

2.1 Atomic structure and periodic table

2.2 Dalton's atomic theory and the modern atomic theory

2.3 Early experiments to characterize the atom

2.4 Makeup of the nucleus

2.5 Electromagnetic radiation (EMR) and atomic spectra

2.6 The quantum mechanical model of the atom

2.7 Electronic configurations and orbital diagrams

2.8 Electronic configurations and the periodic table of the elements

Introduction

Some discoveries in chemistry have had great impact on other disciplines. For instance, chemists have determined the structures and properties of proteins and other biochemical substances, and this knowledge has enabled biologists to study life processes at the molecular level. By the same token, discoveries in other sciences have had dramatic effects on chemistry; considering developments in physics that have profoundly changed the science of chemistry is quite enough. Sometimes discoveries in two or more areas come together at the right time and in a way that changes forever how we look at the world.

In this unit, students will encounter the discoveries in chemistry, starting from early experiments to contemporary science.

The students' textbook has examples, exercises, and activities which are intended to provide supplementary stabilizations to their knowledge of atomic structure and the periodic law.

Start-up Activity

Building blocks are literally the pieces out of which something is built. By now, we are quite sure that atoms are the building blocks of everything around us. For example, in this start-up activity, proteins are built-up of the starting units known as amino acids which, in turn, are composed of C, H, N, and O atoms. Steel contains Fe and C atoms, paper is a substance manufactured in thin sheets from wood fiber, rags, and so on.

2.1 HISTORICAL DEVELOPMENT OF THE ATOMIC NATURE OF SUBSTANCES

Periods allotted: 1 period

Competency

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

- *state briefly the historical development of atomic nature of substances.*

Forward Planning

Atoms are microscopic particles. They can't be seen, touched or weighed directly. This is a very abstract and difficult concept to understand. Even though this is so, convince yourself of the existence of atoms before presenting the lesson to the students. Employ different techniques such as cooperative learning, group discussion, concept map, brain storming and problem solving. You may use questions included in Exercises to acertain how much the students are also convinced of the continuous nature of matter and the existence of the atom.

Take ample preparation time to study the historical development of the atomic nature of substances. Prepare a schematic diagram chronological order showing the historical development of the atomic nature of substances. State clearly the observations and conclusions of these experiments. Set your own contingency plan for helping students to easily understand the processes of the development of atomic nature of substances.

Teaching Aids

No additional teaching aid is suggested.

Subject Matter Presentation

Since most parts of this unit are contained in the textbook of chemistry for Grade 9, assume that students have prior knowledge of atomic structure. Emphasize the sections that are not discussed in detail in Grade 9. Start your discussion by employing *Activity 2.1*, which is supposed to give a clue on how matter is constituted. For instance, in *Activity 2.1***,** table salt is composed of sodium and chlorine atoms, sand from silicon and oxygen atoms; paper from carbon, oxygen and hydrogen atoms.

a. Matter is made up of atoms.

b. Yes, atoms are similar in structure.

c. The idea of atoms is believed to have originated with the Greek philosopher Leucippus and his student Democritus in the 5th century B.C.

d. Atoms are too small to be seen with the naked eye.

e. Atoms are too small to be seen with the naked eye

f. Atoms are divisible into subatomic particles p, n, and e.

Motivate the students to actively participate in *Activity 2.2*. You can then briefly discuss the historical background of the atom. You can explain that the word "atom" is derived from the Greek word "atomos" meaning indivisible. The philosopher Democritus (460 – 370 B.C.) believed that matter was composed of fundamentally indivisible particles called "atomos". According to early philosophers,

- ✓ samples of matter can't be subdivided without limit.
- ✓ tiny, discrete, indestructible units of matter are atoms.
- ✓ atoms have mass.
- ✓ sizes and shapes of atoms determine all material properties.

Assessment

Students working at the minimum requirement level should be able to narrate the historical development of atomic nature of substances. Make them debate the existence or "non-existence" of atoms. Check if the students have acquired the intended knowledge and have a clear understanding of the topic discussed by: Ask questions like "What are the contributions of the ancient Greek philosophers to the contemporary understanding of the nature of matter?" Do not hesitate to record the assessment results on the student-performance list.

2.2 DALTON'S ATOMIC THEORY AND THE MODERN ATOMIC THEORY

Periods allotted:2 periods

Competencies

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

- *state postulates of Dalton's atomic theory;*
- *state postulates of modern atomic theory;*
- *state the laws of conservation of mass, definite proportions, multiple proportions and the basis of each.*

Forward Planning

Take ample preparation time to study Dalton's atomic theory and modern atomic theory. The students are familiar with the concept of the atom and Dalton's atomic theory. Before you begin, prepare questions to check how much the students understand regarding the nature of atoms. After the lesson, comment on Dalton's atomic theory by referring to the summary points of the modern atomic theory.

Teaching Aids

Read *Activity 2.3* and make the materials required ready ahead of time.

Subject Matter Presentation

Postulates of Dalton's Atomic Theory

Scientific laws are the basis for Dalton's atomic theory. Let students discuss this statement and then perform *Activity 2.3*, which will enable students to understand clearly the postulates of Dalton's atomic theory. Let group representatives present their opinions. Ask the individual group members to check whether they all participated in suggesting the ideas being forwarded.

After commenting on the students' work, you may proceed to the lesson on Dalton's atomic theory, using *Activity 2.4*. In this particular activity, the students' response is not supposed to be positive for number 1. The law of conservation of mass can't be violated. If one collects the products of combustion, the mass before and after combustion is exactly the same. 2. Matter can neither be created nor destroyed.

The burned sugar is changed into carbon dioxide gas and water vapor. 3. This partly explains the laws established by Lavoisier and other scientists, which is stated as: *A given compound always contains its constituent elements in certain fixed proportions by mass,* irrespective of its method of preparation or the area found (discovered). 4. A given set of elements can combine in more than one set of proportions, giving rise to different compounds. When these different compounds are compared, the masses of one element that combines with the fixed mass of the second are in the ratio of small whole numbers. This law is known as the law of multiple proportions proposed by John Dalton.

Elaborate the laws, using examples. Exercise 2.1 may be assigned for students to do at home.

Postulates of Modern Atomic Theory

Explain Dalton's atomic theory, and the modifications of his postulates, in contrast to the modern atomic theory. Emphasize that Dalton's atomic theory is considered as the cornerstone of modern physical science. Dalton’s atomic theory had to be modified through time as a result of new discoveries that led to the abandonment of Dalton’s atomic theory with its drawbacks. It is very important to take each of the postulates and show why it failed.

Next, continue to deliver the lesson on modern atomic theory. Let students understand that the modern atomic theory is a modification of Dalton’s atomic theory as a result of recent scientific findings about the structure of the atom, mainly by J. J Thomson, Rutherford and Niels Bohr. Then let the students discuss the postulates of Dalton’s atomic theory that are inconsistent with modern atomic theories, by using *Activity 2.5*. Those which are inconsistent are these postulates:

1. All atoms of a given element are alike in mass and other properties.
2. All matter is composed of extremely small, indivisible particles called atoms.

When one contemplates the early years for the postulates, it may not be difficult to testify that Dalton was farsighted and intelligent. Nobody dares to say that he was wrong, though *Activity 2.6* discusses the limitations of the Dalton’s theory and postulates of modern atomic theory.

1. The main limitations of Dalton's atomic theory are:
   a. It failed to explain why atoms of different elements show different properties like mass, size, valency, etc.
   b. It failed to explain how and why atoms of different elements combine to form compounds (i.e. molecules).
   c. It could not explain Gay Lussac's Law of Combining Volumes.
   d. It failed to explain the nature of the binding force that keeps atoms together in a molecule.
2. The following statements can be taken as a brief summary of modern atomic theory.
   a. Matter is composed of atoms.
   b. Atoms of an element are unique and differ from other atoms of other elements in the number of protons in the atom (i.e., the atomic number).
   c. Chemical compounds are formed when atoms combine in whole-number ratios (law of multiple proportions). A given compound always has the same relative number and types of atoms (law of constant composition).

d. In a chemical reaction, the mass of the reactants (starting materials) equals the mass of the reaction products (law of conservation of mass)

e. Atoms are made of three subatomic particles: protons, neutrons and electrons.

f. Isotopes of elements share the same number of protons, but differ in the number of neutrons in the nucleus. The atomic mass of an element depends on the natural abundance of isotopes of that element.

g. The nucleus of an atom contains protons and neutrons. It is extremely small ($\sim 10^{-15}$ m diameter), enormously dense ($\sim 10^{15}$ g/cm^3) and contains nearly all of the mass of the atom (~99.99%). Neutrons hold the protons together in close proximity within the nucleus with the “strong nuclear force”.

h. Electrons are located outside the nucleus. They are very light (~1/1840th is the mass of a proton or neutron) very fast-moving particles, and occupy the vast majority of the space in the atom.

i. New quantum physics models had to be developed to explain the observed behavior of both electrons and of light. Like light, electrons have a dual nature; they behave both as particles and as waves. Like light, the energy of the electron is quantized; energy exists in small “packets” called quanta.

j. The electron has energy because of its motion relative to the nucleus. Energy of the electron is quantized: the electron is restricted to only certain energy levels and cannot exist at intermediate levels. The wave nature of electrons means that their behavior can only be described in terms of probability. These probability maps for electrons with specific energy levels are called orbitals. Each orbital describes the behavior of two electrons, and these must have opposite spin.

k. The ground state of an atom is when all the lower energy level orbitals are filled with electrons. An excited state occurs when one or more electrons absorb exactly the right amount of energy to jump to a higher energy state orbital (an unstable circumstance). As an electron returns to the ground state, it emits light of characteristic wavelength, corresponding to the difference in the energy levels between the excited state and the ground state. This gives each atom a unique emission spectrum.

l. An element can be characterized by its electron configuration. This configuration describes how the electrons populate orbitals at the various energy levels; each atom has a unique ground-state configuration. Electron configuration, based on the quantum theory of the atom, explains the structure of the periodic table, the chemicals and physical properties of elements, chemical reactivity and chemical bonding.

m. The quantum theory of the atom is the most accurate and most successful scientific theory yet developed. It is highly predictive, and all predictions subjected to experiment have proved correct. It has never been wrong yet.

How the Modern Theory Explains The Mass Laws

let the students discuss *Activity 2.7*, which helps them to understand how the modern atomic theory explains the mass laws.

Assessment

Students working at minimum requirement level should be able state postulates of Dalton's and Modern atomic theories. Evaluate if the students could clearly distinguish between these two theories. In order to do so, you may use different instruments of assessment such as class works, home works, assignments, quizzes or tests.

Answers to Exercise 2.1

1. i. Compounds are formed when atoms of different elements unite in fixed proportions.

   ii. No atoms are created, destroyed, or broken apart in a chemical reaction.

2. I. C II. A III. B

3. A given compound always contains its constituent elements in certain fixed proportions by mass.

4. a. 1.176 g b. the law of definite proportions

   c. mentioned above in question number 3: A given compound always has the same relative number and types of atoms (a given compound always has the same elements in a certain fixed proportions by mass).

5. a. 1.43 g O / 1.00 g N; 2.286 g O / 1.00 g N; 2.857 g O / 1.00 g N

   b. these masses of oxygen per gram of nitrogen are in the ratio of 2 : 4 : 5 and thus obey the law of multiple proportions.

2.3 EARLY EXPERIMENTS TO CHARACTERIZE THE ATOM

Periods Allotted: 3 periods

Competencies

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

- *discuss the discovery of the electron;*
- *describe the properties of cathode rays;*
- *define the terms: radioactivity, radioactive decay and radio-isotope;*
- *describe the common types of radioactive emissions;*
- *discuss the alpha scattering experiment;*
- *describe the major contribution of experiments by Thomson, Millikan and Rutherford concerning atomic structure.*

Forward Planning

Take ample preparation time study the early experiments to characterize the atom. You are supposed to read and understand the experiments that led to the discovery of the electrons and the nucleus. Sketch schematic diagrams of the cathode – ray tube, the apparatus Robert Millikan used to determine the charge on the electron and the alpha-particle scattering experiment by Rutherford. Clearly understand the observations and conclusions of these experiments. Make the necessary preparations to explain radioactivity and the common types of radioactive emissions.

Teaching Aids

Tubes, copper wires, etc. to prepare a model and flipcharts are required.

Subject Matter Presentation

Discovery of the Electron

You may apply teaching-learning methods such as group discussion, concept map, and brain storming. Start this lesson by asking students to discuss, in groups, the discovery of electrons and the properties of cathode rays, based on their Grade 9 Chemistry. Let the group representatives present the opinions of the group members.

In teaching the discovery made by J. J. Thomson, you may use ordinary materials around you, such as transparent pen tubes and copper wires, to make a simple model of a cathode ray tube. In your explanation, try to help students visualize how electric discharges follow a gentle path from one electrode to another as they flow in the discharge tube. Let the students read the text note in groups and motivate them to ask questions related to the content.

Make sure that students understand the following properties of cathode rays:

- The “ rays” travel in a straight line.
- The “ rays” are attracted towards the positive electrode.
- The “rays” are not directly visible but can be detected by their ability to cause other materials to glow.
- They deflect by the influence of magnetic or electric fields.
- The same properties are exhibited by the “cathode rays” produced by cathodes of different materials.
- A metal plate in the path of the “cathode rays” acquires a negative charge.

Discuss with the students that these observations indicated that the cathode ray radiation was composed of negatively charged particles now known as **electrons**.

Articulate the fact that J. J. Thomson could only determine the charge-to-mass ratio for the electron, but was unable to determine the mass of the electron. He measured the charge-to-mass ratio for a stream of electrons (using a cathode ray tube apparatus) and found it to be 1.76×10^{8} coulombs/gram. Robert Millikan was able to successfully measure the charge on a single electron. You can use the schematic diagram of the Milikan oil drop experiment to explain how the charge on an electron was determined. The charge of an electron was determined to be 1.60×10^{-19} coulombs. Thus, the mass of a single electron was determined to be:

$$\left(\frac{1g}{1.76\times10^{8}\text{ colombs}}\right)\times\left(1.60\times10^{-19}\text{colombs}\right)=\left(9.10\times10^{-28}\text{grams}\right)$$

Compare this value with the accepted current value (9.10939×10^{-28} grams)

Radioactivity and Discovery of Nucleus

Teaching-learning methods such as group discussion, concept map, and brain storming may be applied. After realizing that students have understood the previous lesson, continue teaching radioactivity. Discuss radioactivity briefly and mention the three different kinds of particles emitted by radioactive decay. The history of findings by Antoine Henri Becquerel is an interesting story to narrate. Here, again use diagrams prepared on flipcharts to illustrate the radiation emissions. Involve your students in discussions using *Activity 2.8* and *2.9*, which enable students to understand the types and magnetic properties of radioactive emissions. *Activity 2.9* is about Rutherford’s alpha-particle scattering experiment.

Types of radiation	***Greek Letters***	***Mass Number***	***Charge***	***Deflected***
Alpha particles	α	4	+ 2	Towards the negative plate
Beta particles	β	0	−1	Towards the positive plate
Gama rays	γ	0	0	Undeflected

Rutherford explained the result of the α-particle scattering experiment in terms of a new model for the atom, according to which:

- Most of the space in the atom is empty. This explains why the majority of α particles pass through the gold foil.
- The atom has a positive charge, all of which is concentrated in the nucleus, which is a dense central core within the atom. The heavy positively charged core is responsible for the deflections and the bouncing back of the α particles.

Then proceed to discuss Rutherford's explanation of the α-scattering experiment. Use Figure 2.3 drawn on flipcharts. Let students notice that this lead to the discovery of the nucleus and to a new model of the atom. Continue on by briefly discussing the discovery of the neutron. In the summary, help the students to appreciate the experiments to characterize the atom that were made by J. J. Thompson, Ernest Rutherford and James Chadwick.

Assessment

At the end of each part of the lesson, ask short-answer questions. See whether the lesson objective has been achieved or not. Evaluate students' understanding of what they read from the text.

Assess each student's work to determine whether the students have achieved the minimum required level or not. Students working at the minimum requirement level will be able to:

- discuss the early experiments to characterize the make-up of the atom;
- list the three most important sub-atomic particles, their charges and relative masses;
- compare and contrast Dalton's and the modern atomic theories;
- describe the properties of cathode rays; and
- define radioactivity, radioactive decay and radioisotope; and describe the common types of radioactive emissions.

Appreciate students working above the minimum requirement level and recognize their achievements. Encourage them to continue working hard and not become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class. Give them extra attention in class and additional lesson time during breaks or at the end of the day.

Additional Questions

*1. The mass-to-charge ratio of the proton is found to be 1.044×10^{-8} kg / C. The charge on the proton is 1.602×10^{-19} C. Calculate the mass of the proton.

2. An oil-drop experiment results in the following charges on oil droplets:

6.4×10^{-19} C, 3.2×10^{-19} C, and 4.8×10^{-19} C. What value for the electronic charge can be deduced from this experiment?

* 3. What is a cathode ray? How was the nature of cathode rays determined?

4. Describe J. J. Thomson's experiment that determined the mass-to-charge ratio of the electron.

5. Describe Millikan's oil-drop experiment. How did Millikan determine the charge on an electron?

* 6. What is radioactivity? How was it discovered?

7. How do gamma rays differ from X-rays? How are the two kinds of rays similar?

8. How did the discovery of radioactivity cause Dalton's atomic theory to be modified?

Answers to Additional Questions

1. 1.672×10^{-27} kg
2. All values are integral multiples of 1.602×10^{-19} C.
3. Cathode rays (also called an electron-beam or e-beam) are streams of electrons observed in vacuum tubes which are evacuated glass tubes that are equipped with at least two metal electrodes to which a voltage is applied, a cathode or negative electrode and an anode or positive electrode.
4. In 1897, British physicist J. J. Thomson showed the rays were composed of a previously unknown negatively charged particle, which was later named the electron. Thomson's most significant experiment was the one which was used to calculate the ratio of the mass of an electron to its charge (m_e/e). Cathode ray tubes (CRTs) create the image in a classic television set. Cathode rays are so named because they are emitted by the negative electrode, or cathode. This research led to further experiments

by Bohr and Rutherford, leading to an understanding of the structure of the atom.

5. The oil-drop experiment was an experiment performed by Robert Millikan and Harvey Fletcher in 1909 to measure the elementary electric charge (the charge of the electron). Using a known electric field, Millikan and Fletcher could determine the charge on oil droplets. By repeating the experiment for many droplets, they confirmed that the charges were all multiples of some fundamental value, and calculated it to be 1.5924×10^{-19} C, which is within one percent of the currently accepted value of $1.602176487\times10^{-19}$ C. They proposed that this was the charge of a single electron.

6. Radioactivity is the spontaneous emission of energy from unstable atoms. Atoms are found in all natural matter. There are stable atoms, which remain the same forever, and unstable atoms, which break down or 'decay' into new atoms. This phenomenon is observed in the heavy elements, like uranium, and unstable isotopes, like carbon-14. These unstable atoms are said to be 'radioactive', because they emit radiation from the nucleus as they decay.

   Radioactivity was first discovered in 1896 by the French scientist Henri Becquerel, after whom the SI unit for radiation, the Becquerel, is named. Becquerel discovered that uranium salts were able to blacken a photographic plate placed in the dark. Subsequent experiments distinguished three distinct types of radiation: alpha particles, beta particles, and gamma rays. These are positively charged, negatively charged, and neutral, respectively.

7. Some of the differences between X-rays and gamma rays are listed below. The key difference is the source: X-rays are emitted by the electrons outside the nucleus, and gamma rays are emitted by the excited nucleus itself. Another difference lies in their frequencies. X-rays' frequencies are lower than those of gamma rays. Their wavelengths also vary. The gamma rays' wavelength is smaller than that of the X-rays'. Gamma-ray photons have the highest energy in the EMR spectrum, and their waves have the shortest wavelength. Gamma rays are a lot more dangerous and hazardous to human health than X-rays. Moreover, gamma rays are the highly penetrating and highly energetic ionizing radiation. On prolonged exposure to living beings, they can cause cancer. Since their wavelength is very small, they have an ability to penetrate through any gap, even if it is a subatomic gap.

   Gamma rays are sometimes produced along with other types of radiation like alpha and beta. However this is not the case with X-rays.

   The similarities of gamma rays and X-rays lie in that they are both types (forms) of electromagnetic radiation travelling at the same speed.

8. Dalton believed that the atom was indivisible, although the discovery of radioactivity cast doubt in the minds of many scientists. At the same time it was generally believed that electric charge, like mass, was infinitely divisible.

2.4 MAKEUP OF THE NUCLEUS

Periods allotted: 2 periods

Competencies

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

- *describe make-up of the nucleus;*
- *define atomic mass;*
- *define isotope;*
- *calculate the relative atomic mass (atomic mass) of naturally occurring isotopic elements.*

Forward Planning

Be sure that students are well acquainted with the properties of n, p, e, atomic mass, and isotopes from their Grade 9 chemistry lessons.

The complete understanding of the natures and properties of protons and neutrons are extremely important prior to presenting the topic to the students. Advance preparations on atomic number, mass number, relative atomic mass (RAM) and isotopes are recommended. Take ample preparation time to study the makeup of the nucleus.

Teaching Aid

No additional teaching aids are suggested.

Subject Matter Presentation

Constituents of the Nucleus

Teaching-learning methods such as group discussion, gaped lecture may be applied. After a short review of the previous topic, introduce your current topic. Let students perform *Activity 2.10* during the discussion session. This will help you to evaluate their understanding. It is intended to refresh the students' prior knowledge. As you check the student groups to facilitate and assess students' discussion, identify the problems faced by students in answering the questions. Harmonize their responses with the facts presented in the textbook.

Symbol	*Atomic number*	*Number of neutrons*	*Mass number*	*Number of electrons*
${}^{12}_{6}C^{+}$	6	6	12	5
S^{2-}	16	16	32	18
Al	13	14	27	13
Pb	82	126	208	82

Atomic Mass and Isotopes

You can now go to discussing atomic mass and isotopes. In the meantime, you can engage students in performing *Activity 2.11*, which is planned to relate their Grade 9 knowledge with that of Grade 11. Support your explanation with the examples given, and have the students work on Exercise 2.4 as homework. The suggested answers for *Activity 2.11* are given below.

1. The answer is a plain no! Because hydrogen (^1H) does not exhibit any neutron.
2. No; atoms of different elements have different numbers of electrons.
3. Usually an atom has an equal number of protons and electrons.
4. Chemical properties depend only on the number of electrons.
5. Because atomic masses are the average of the isotopes, and, most because averages are not whole numbers.
6. The relative atomic mass of Zn is 65.39, and that of S is 32.066, which is about half of that of Zn.
7. Nitrogen-14 is abundant; because the relative atomic mass of nitrogen is near 14.

Assessment

Check whether students participate actively in discussions. Don't forget to record students' performance in the performance list as you evaluate their homework.

Students working at the minimum requirement level will be able to:

- describe the constituents of the nucleus;
- define atomic mass, mass number, and isotope
- trace the difference between mass number and atomic mass; and
- calculate the relative atomic masses of naturally occurring isotopic elements.

Praise students working above the minimum requirement level and recognize their achievements. Encourage them to continue working hard and not become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class.

Additional Questions

* 1. What is an atom?

* 2. Does a proton have more mass than a neutron?

3. Does a proton have more mass than an electron?

* 4. What is atomic number?

5. What is an isotope?

* 6. What is atomic structure?

7. Who discovered the atomic structure?

Answers to Additional Questions

1 An atom is the smallest piece of a substance that can exist.

2 Protons and neutrons have about the same mass.

3 An electron is very much smaller. It is about 1÷2000 times the size of a proton, although it has an equal and opposite electrical charge.

4 The number of protons in the nucleus is called the atomic number. The atomic number tells you what the element is. The atomic number of an element never changes.

5 Atoms of the same element can have different numbers of neutrons. These are called isotopes.

6 Atomic structure is the way in which atoms are arranged in a molecule with specific details about protons, electron, and neutrons.

7 Niels Bohr and Ernest Rutherford are credited with discovering the atomic structure. Although Bohr was the one to publish the findings describing atomic structure in 1913, his work would not have been possible without Rutherford's gold foil experiment that discredited earlier theories.

Answers to Exercise 2.2

1. 116
2. Mg: 12 protons, 12 neutrons Sr: 38 protons, 50 neutrons
3. Because they have the same number of electrons
4. Element X is Cr; ${}^{52}_{24}\text{Cr}$
5. ${}^{63}_{29}\text{Cu} = 69.1\ \%$, ${}^{65}_{29}\text{Cu} = 30.9\ \%$
6. ${}^{35}_{17}\text{Cl} = 73.5\ \%$, ${}^{37}_{17}\text{Cl} = 26.5\ \%$
7. 12.01 m_u

2.5 ELECTROMAGNETIC RADIATION (EMR) AND ATOMIC SPECTRA

Periods allotted: 9 periods

Competencies

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

- *characterize electromagnetic radiation (EMR) in terms of wave length, frequency and speed;*
- *calculate the wave length and frequency of the EMR;*
- *explain that light has both wave and particle natures;*
- *define photon as a unit of light energy;*
- *explain how photon theory explains the photoelectric effect and the relation between photons absorbed and electrons released;*
- *explain that emission spectra of atoms consist of series of lines;*
- *state Bohr's assumption of energy of electron in the hydrogen atom;*
- *explain that the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron;*
- *explain that atoms emit or absorb energy when they undergo transition from one state to another;*
- *explain the shortcomings of Bohr theory;*
- *calculate the radius of electron orbits, the electron velocity and the energy of the electron, using the Bohr model.*

Forward Planning

Take ample preparation time to study electromagnetic radiation (EMR) and atomic spectra. Bear in mind that sections 2.5 and 2.6 are a little bit challenging for the students. Therefore, you must read and try to thoroughly understand the concepts of the topic. Combine what there is in the textbook of Grade 11 with other relevant references. Get prepared, so that you will be able to present an excellent presentation.

Teaching Aids

Diagrams prepared on flipcharts for electromagnetic radiation spectrum, and Bohr's model for hydrogen atom are required

Subject Matter Presentation

Electromagnetic Radiation

After introducing the topic, let students perform *Activity 2.12* which engages the power of the students' minds in understanding the nature and position of electrons in an atom. It uses the problem-solving method, which helps to identify, in advance areas of difficulty by thinking of questions which will help students in the following topics. The suggested answers for *Activity 2.12* are given below.

1. Predicting the exact location of an electron, according to quantum theory, is only a matter of probability. The Heisenberg uncertainty principle states that it is impossible to know both the location and direction of an electron. You can know the location, but then the direction of the electron will change. If you find the direction, the location will change.
2. Electrons have dual natures – particle and wave natures. The atomic models are the manifestations of its particle nature, and the spectral lines indicate its wave nature.
3. The electrostatic attraction ($F= Kq_1q_2/r^2$) between the electron and the nucleus, which tends to pull the electron towards the nucleus, must be equal to the centrifugal force ($F = mv^2 /r$), which tends to throw the electron out of its orbit.
4. It is possible to calculate the velocity of an electron using, for example, the de Broglie equation. Electrons that have a wave length of 12 pm have a velocity of 6.1×10^7 m/s.

After the students present their opinions, continue your discussion on properties of electromagnetic radiation. Stress the main ideas of electromagnetic radiation (EMR) – wavelength, frequency, amplitude, and regions of electromagnetic radiation. Select appropriate ways of demonstrating the wave like and particle-like properties of electromagnetic radiation. Explain the core idea of quantization. Tell the students that much of our understanding of the structure of the atom comes from the study of the interaction of light with matter. Explain that visible light is only a small part of the electromagnetic spectrum. Try to make the students understand that each region of electromagnetic radiation is characterized by its own frequency (ν) and wavelength (λ), and that all electromagnetic radiation travels at the same speed. Explain the characteristic parameters of a wave, such as wavelength, amplitude and frequency and their relationships using Example 2.3 and Exercise 2.3.

The Quantum Theory and Photon

Use figures prepared on flipcharts as usual to discuss the quantum theory. Support your discussion, using examples, and assign students to work on Exercise 2.4. Then you can proceed to explain the quantum theory of radiation, using the discussion method. Start class discussions on the quantization of energy by mentioning that the classical laws of physics said that radiation was continuous, and that, therefore, any energy value is possible. Try to help students visualize the concept of *quanta* and *quantum.* You may ask your students whether radiation is countable or not. Describe the analogies given in the textbook in relation to the quantum theory of photons.

Help them notice that radiant energy is quantized and can only be emitted in discrete amounts called quanta. A quantum of radiation is a photon. Support the discussion by giving relevant examples.

Move directly to the lesson on photoelectric effect. As an activity, let the students read the explanation given in their textbook and understand the core idea of the analogy given for the concept of photoelectric effect. Let them try to represent what they visualize from their reading, using some kind of sketch or diagram. In the discussion that follows try to refine their mental picture while explaining the concept. Support the discussion, using Example 2.5.

Emphasize the following key points when you discuss the photoelectric-effect experiment.

- Light shining onto the surface of a metal is capable of ejecting electrons from that metal.
- The electrons that are emitted come off at energies that depend up on the frequency of the light, and not on its intensity.

Help the students to clearly understand that the photoelectric effect can be explained if one assumes that light is composed of photons of energy, $h\nu$. Albert Einstein suggested that a photon could give up its energy to an electron in the metal. Unless students understand, the dual nature of light, the photoelectric-effect experiment is difficult to explain.

Atomic Spectra

Continue by teaching atomic spectra after ensuring the students' understanding of the photoelectric effect. Figure 2.7, drawn on a flipchart, is helpful. Explain that an analytical technique called emission spectroscopy is based on an analysis of the light emitted when an element is strongly heated or energized in an electric spark or a gas discharge tube. Help the students to have a clear idea about the light emitted by excited atoms and that it is dispersed into line spectra. State that every element has a unique emission spectrum which serves as an 'atomic finger print' and can be used to identify the element.

The Bohr Model of the Hydrogen Atom

Various teaching-learning methods such as group discussion, concept map, brain storming, and analogy may be applied. Proceed to teaching the Bohr model of the hydrogen atom. Let students discuss the Bohr model of an atom and how he developed it by combining ideas from classical physics and the new quantum theory. After commenting on the students' responses, let students read Bohr's postulate of his model. Ask them to present their visualizations, using diagrams. Make some corrections to the students' work and them present a clear picture of the model.

Ensure that the students clearly understand that Bohr's model is based on the findings described in the preceding topic. Bohr made clear the assumption that the energy content of a system is not continuous. Variations can be visualized in terms of energy levels. Energy absorbed or emitted involves the transition of a component of the system between energy levels.

Use the analogy of imagining the various energy levels as rungs on a ladder. One can step on the rung either one at a time, or more if he/she is daring enough to do so. Nevertheless, one cannot stand at a point between two rungs. This analogy helps the students to understand that an electron's energy cannot be in between two energy levels.

Ensure that the quantum theory that explains the origin of spectral lines is applied during the discussion of the Bohr model of the hydrogen atom. Show the students that, when atoms are subject to high temperature or electric discharges, they emit electromagnetic radiation, which consists of characteristic frequencies. By subjecting the atoms of a given element to very high temperatures or passing an electric discharge through a gas composed of these atoms, the atoms become excited, as stated, under atomic spectra. This shows that once electrons within the atoms are raised to higher energy levels, they can return to their ground state by losing this increased energy. This can be achieved by emitting EMR (or light).

The frequency of the emitted radiation will be determined by the energy of the "jump" made by the electron. The electron could return to the ground state in a single transition, or it could return by a series of transitions. The analogy is needed here. You can get down off a ladder either by stepping down to one rung at a time or by making a jump of two or three rungs. Each jump will be accompanied by a release of radiation, each being characteristic of the magnitude of the jump. Because there are a number of possible routes back to the ground state, what we observe is a series of lines, each line corresponding to a jump from a higher energy level to a lower one.

Help the students to recognize the physical meaning of *quantum number* as related to the size of the radius of the orbit assumed by the electrons possessing certain amounts of energy by using analogy. Give them enough time to read the contents of their textbook and to understand the meaning of the mathematical formulae given to describe Bohr's model. After summarizing, assign them Exercise 2.6. Finally, explain the limitations of the Bohr theory. As you discuss de Broglie's hypothesis, involve the students in discussion, reminding them of the relevant contents related to electromagnetic radiation.

Assessment

You are required to assess students' active involvement in each of the activities. Note that this subject is quite difficult. Don't forget to appreciate students working above the minimum requirement level. You are required to assist those working below the minimum requirement level.

Students working at the minimum requirement level will be able to:

- characterize electromagnetic radiation (EMR) in terms of wave length, frequency and speed;
- calculate the wave length and frequency of the EMR;
- explain that light has both wave and particle natures;
- explain that the emission spectra of atoms consist of series of lines;
- state Bohr's assumption of energy of electron in the hydrogen atom;
- calculate the radius of an electron orbit, the electron velocity and the energy of the electron, using the Bohr model;
- explain that atoms emit or absorb energy when they undergo transition from one state to another; and
- explain the shortcomings of the Bohr theory.

Additional Questions

* 1. Radio station KF operates at a frequency of 992 KHz on the AM band. What is the wavelength, in meters, of the radio waves?

* 2. Calculate the energy, in joules, of a photon of red light that has a frequency of $3.73 \times 10^{14}\ s^{-1}$

3. A microwave oven operates with radiation having a wave length of 1.20 cm. Calculate the energy of this radiation, in joules per photon and in kilojoules per mole of photons.

4. Calculate the frequency, in s^{-1} of the electromagnetic radiation emitted by a hydrogen atom, when its electron drops from (a) the n=3 level to the n=2 level and (b) from the n=4 level to the n=1 level.

* 5. Electrons have an associated wavelength of 84.4 nm. At what speed, in meters per second, are the electrons traveling?

* 6. Calculate the wavelength, in nanometers, associated with a proton traveling at a speed of 2.55×10^6m /s. Use 1.67×10^{-27} kg as the proton mass.

Answers to Additional Questions

1. 302 m
2. 2.47×10^{-19} J
3. 1.66×10^{-23} J/photon; 0.0100 kJ/mol
4. a. 4.568×10^{14} s^{-1} b. 3.083×10^{15} s^{-1}
5. 8.62×10^3 m/s
6. 1.56×10^{-4}nm

Answers to Exercise 2.3

1. 415 nm
2. 2.94 m, 2.94×10^9 nm, 2.94×10^{10} Å

Answers to Exercise 2.4

1. Both are electromagnetic waves with similar poles and charges (north/south, positive/ negative) and the same constant speed.
2. For a given frequency of incident radiation, the rate at which photoelectrons are ejected is directly proportional to the intensity of the incident light. An increase in the intensity of the incident beam increases the magnitude of the photoelectric current.
3. $$\Delta E = h\nu = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\,Js\times3\times10^{8}\,ms^{-1}}{350\times10^{-9}\,m} = 5.678\times10^{-19}\,J$$

$$h\nu_o = 6.626\times10^{-34}\,Js\times5.46\times10^{14}\,s^{-1} = 3.617\times10^{-19}\,J$$

$$\Delta E = h\nu = h\nu_o + E_k$$

$$E_k = h\nu - h\nu_o = 5.678\times10^{-19}\,J - 3.617\times10^{-19}\,J = 2.06\times10^{-19}\,J$$

$$E_k = \frac{1}{2}m_e v^2\,;\ v = \sqrt{\frac{2E_k}{m_e}} = 6.7\times10^5 m/s$$

4. $$\frac{hc}{\lambda} = \frac{6.626\times10^{-34} Js\times3\times10^8 ms^{-1}}{632.8\times10^{-9} m} = 3.14\times10^{-19} J/photon$$

$$\Delta\text{E} =\ 3.14\times10^{-19} J/photon \times6.02\times10^{23} photon/mol = 1.89\times10^2 kJ/mol$$

5. In = 4.4×10^{-19} J; Th = 3.7×10^{-19} J

Answers to Exercise 2.5

1. Yes, the predicted line spectra are accurate. The energies could be predicted from $\text{E}_\text{n} = \frac{(-\text{Z}^2)(2.18\times10^{-19})}{\text{n}^2}$, where Z is the atomic number for the atom or ion. The energy levels for the Be^{3+} will be greater, by a factor of 16, than those for the hydrogen atom. This means that the pattern of lines will be similar, but at different wavelengths.

2. $$\Delta E = R_H\left(\frac{1}{n_f^2}-\frac{1}{n_i^2}\right) = -2.18\times10^{-18} J\left(\frac{1}{3^2}-\frac{1}{2^2}\right) = 3.03\times10^{-19} J$$

$$\Delta E = h\nu = \frac{hc}{\lambda};\ \lambda = \frac{hc}{\Delta E} = \frac{6.626\times10^{-34} Js\times3\times10^8 ms^{-1}}{3.03\times10^{-19} J} = 6.56\times10^{-7} m = 656\ nm$$

3. $$\Delta E = R_H\left(\frac{1}{n_f^2}-\frac{1}{n_i^2}\right) = -2.18\times10^{-18} J\left(\frac{1}{5^2}-\frac{1}{2^2}\right) = 4.58\times10^{-19} J$$

$$\lambda = \frac{hc}{\Delta E} = \frac{6.626\times10^{-34} Js\times3\times10^8 ms^{-1}}{4.58\times10^{-19} J} = 4.34\times10^{-7} m = 434\ nm$$

4. $$\Delta E = -2.18\times10^{-18} J\left(\frac{1}{5^2}-\frac{1}{2^2}\right) = 4.58\times10^{-19} J/atom$$

For 1 mole of atoms,

$$\Delta E = 4.58\times10^{-19} J/atom \times6.02\times10^{23} atom/mol = 2.76\times10^5 J/mol = 276\ kJ/mol$$

276 kJ; no it is not sufficient

5. The question on the student text needs modification as follows:

“Three lines in the hydrogen atom spectrum result from transition to n_f= 2 (visible series). These are

a. 4340.5Å b. 4861.3 Å c. 6562.8 Å

The other two result from transitions to n_f =1 (d) and n_f =3(e):

d. 1217 Å e. 10938.0 Å"

Answer: Using the formula $\Delta E = \frac{hc}{\lambda} = R_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$.

a. $5 \rightarrow 2$ b. $4 \rightarrow 2$ c. $3 \rightarrow 2$

d. $2 \rightarrow 1$ e. $6 \rightarrow 3$

Answers to Exercise 2.6

1. Using the formula $\lambda = \frac{h}{mv}$, $\lambda = 0.123$ nm

2. $\Delta E = R_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$; insert $n_f = \infty$ and $n_i = 1$, $\Delta E = 2.18 \times 10^{-18}$ J

3. 434nm

2.6 THE QUANTUM MECHANICAL MODEL OF THE ATOM

Periods allotted: 5 periods

Competencies

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

- *state the Heisenberg uncertainty principle;*
- *describe the significance of electron probability distribution;*
- *explain the quantum numbers n, l, m_l and m_s;*
- *write all possible sets of quantum numbers of electrons in an atom;*
- *describe the shapes of orbitals designated by s, p, and d.*

Forward Planning

Take sufficient time to study the quantum mechanical model of the atom. To have sufficient knowledge and become prepared to explain the subject matter, you must obtain and read a lot more reference material than the grade-11 chemistry textbook. Clearly, the quantum mechanical model is highly mathematical, and its full treatment is beyond the scope of the students. After reading and understanding the significance of the four quantum numbers, practice writing the possible combinations of quantum numbers that define an electron in a given shell and subshell.

Teaching Aids

Diagrams prepared on flipcharts, showing electron probability distribution curves and charts showing the shapes of s, p and d orbitals.

Subject Matter Presentation

Start discussing this topic by asking the students to discuss *Activity 2.13*, which gives bird's-eye-view information on the sophisticated mathematical description of atomic structure based on the wave properties of subatomic particles. By now, students should regard matter as consisting of particles and understand the possibility that matter could, under proper circumstances, behave as waves.

Regarding *Activity 2.13*, macroscopic objects do exhibit a wavelike motion, but the wavelength is too small for humans to perceive. Though it is known that electrons possess particle nature, they are microscopic particles which move with a very high speed in which it is very difficult to trace their exact locations. This is the reason why people talk about the high probability region of finding an electron. A wave is associated with any object, even a moving elephant, but one can't observe it for the reasons as same mentioned for macroscopic objects.

The Heisenberg's Principle

Start the discussion on the quantum mechanical model of the atom by presenting the idea of the Heisenberg uncertainty principle. The contribution of this principle to the modification of the atomic model should be stressed. Note that *Activity 2.12* relates to Heisenberg's uncertainty principle. Discuss it with the students, clearly explaining that we can't simultaneously know exactly where a tiny particle is and exactly how it is moving.

Quantum Numbers

Now, mention that we must have a model of the atom where "probability" is important-High probability of finding an electron within a certain volume of space. Here it seems that an analogy is important. Assume that you are explaining the location of your residence to someone coming from outside of Ethiopia**.** What you are supposed to explain is probably: the subcity, woreda , kebele and house number. Obviously, with that information the incoming person can easily find your residence. Similarly, continue to explain the necessity to use the four quantum numbers to describe an orbital and the high probability region for finding an electron within this orbital. Discuss these four quantum numbers, taking the following points into consideration:

- The principal quantum number (n)
    - determines the size of an orbital (i.e., larger n is larger orbital).
    - determines the energy of the orbital (i.e., larger n is higher energy orbital).
    - can take integral values n = 1, 2, 3------ ∞.
    - all electrons in an atom with the same value of n are said to belong to the same shell.
- The angular momentum, azimutal or subsidiary quantum number (ℓ)
    - designates orbital energy shape within a shell.
    - affects orbital energy/larger ℓ is higher in energy/.
    - all electrons in an atom with the same value of ℓ are said to belong to the same subshell. Only integer values between 0 and n – 1 are allowed.
- The magnetic quantum number mℓ
    - determines the orientation /direction/ of orbitals.
    - the number of m_ℓ values within a subshell, is the number of orbitals within a subshell.
- The spin quantum number (m_s)
    - several experimental observations can be explained by treating electrons as though they were spinning.
    - spin makes an electron behave like a tiny magnet.

Support your discussion by using the examples given, and have the students work on Exercise 2.7 as homework. Then you may move on to discuss shapes of atomic orbitals, using diagrams.

Shapes of Atomic Orbitals

Start the discussion by asking a question like: “What do the regions of high electron probabilities described by atomic orbitals look like?” Of course, one can’t see these regions, but the mathematical forms of the orbitals can be translated into geometric pictures. Informing students of the geometric shapes of orbitals at this level is quite enough.

Assessment

In addition to the exercise given, ask short-answer questions at the end of each session or lesson period to assess the understanding of students. Record their performances on the students' performance list.

Students working at the minimum requirement level will be able to:

- state the Heisenberg uncertainty principle and describe the significance of electron probability distribution;
- explain the quantum numbers n, l, m_l and m_s; and
- clearly undestand the shapes of orbitals designated by s, p, and d.

Additional Questions

* 1. Give the subshell notation for each of the following sets of quantum numbers.

a. $n = 3$ and $l = 2$ b. $n = 2$ and $l = 0$

c. $n = 4$ and $l = 1$ c. $n = 4$ and $l = 3$

* 2. How many sublevels are there in an energy level with each of the following values of n?

a. $n = 3$ b. $n = 2$ c. $n = 4$

* 3. If n=5, what are the possible values of l? If $l = 3$, what are the possible values of m_l ?

4. Consider the electronic structure of an atom:

a. What are the n, l, and *ml* quantum numbers corresponding to the 3s orbital?

b. List all the possible quantum number values for an orbital in the 5f subshell.

c. In what specific subshell will an electron having the quantum numbers n=3, l=1 and m_l=-1 be found?

Answers to Additional Questions

1. a. 3d b. 2s c. 4p d. 4f

2. a. 3 sublevels b. 2 sublevels c. 4 sublevels

3. If $n = 5$, $l = 0,1,2,3,4$

If $l = 3$, $m_l = -3, -2, -1, 0, +1, +2, +3$

4. a. for 3s, $n = 3, l = 0, m_l = 0;$

   b. for 5f, $n = 5$, $l = 3$, $m_l = -3, -2, -1, 0, +1, +2, +3$

   c. 3p

Answers to Exercise 2.7

1. a. 3d b. 2s c. 4p d. 4f

2. a and d are not permissible because

   a. when $l = 1, m_l = -1,0,+1$ but not +2

   d. Again, n can't be 0, and the values of n and l can't be equal. l is $n - 1$.

3. a. $n = 3$, $l = 0$, $m_l = 0$

   b. $n = 5$, $l = 3$, $m_l = -3,-2,-1,0,+1,+2,+3$

   c. $3p_x$

4. The spin quantum number, magnetic quantum number.

2.7 ELECTRONIC CONFIGURATIONS AND ORBITAL DIAGRAMS

Periods allotted: 2 periods

Competencies

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

- *explain the Aufbau principle;*
- *explain the Pauli exclusion principle;*
- *explain Hund's rule;*
- *write the ground state electronic configurations of multi-electron atoms.*

Forward Planning

Take sufficient time to study the electronic configurations and orbital diagrams. Note that students of Grade 11 have prior knowledge of electronic configuration.

Teaching Aids

No additional teaching aids are suggested.

Subject Matter Presentation

In your presentation, guide the students in applying Aufbau's principle, the Pauli exclusion principle, and Hund's rule to write the electronic configuration of atoms along with their quantum numbers. Demonstrate the Pauli exclusion principle by using a combination of the four quantum numbers.

Then continue by introducing the other rules / principles/ by giving examples, and let the students notice that there are exceptions to the Aufbau principle. Use Activities 2.14-2.18, which are thought to motivate students into discussions.

The suggested answers for *Activity 2.14* are given below.

1. Because an atomic orbital is a mathematical function that describes the wave-like behavior of either one electron or a pair of electrons. They are the possible quantum states of an individual electron in the collection of electrons around a single atom.
2. The spin of an electron is a magnetic property, not an actual rotation like a magnet with its south and north poles: Electrons which are paired have opposite magnetic properties, which leads to opposing spins.
3. It reduces electron-electron repulsion energy. I.e., there is a total lowering of energy – Hund's rule of degenerate orbitals.
4. The greatest total spin state usually makes the resulting atom more stable – Hund's rule of maximum multiplicity.

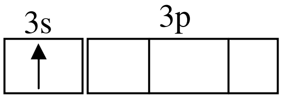
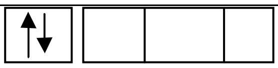
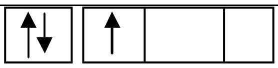
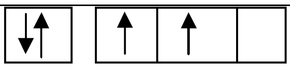
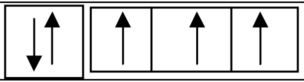
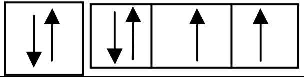
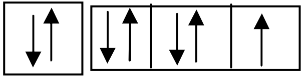
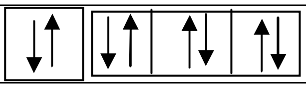
The suggested answers for *Activity 2.15* are given below.

Can be used to identify the three types of quantum numbers (n, l, and m_l) that directly relate to the energy, orientation and shape of the orbitals in which the values may vary for various electrons of an atom. However, m_s specifies only the orientation of the spin axis of an electron, and an electron can spin in only one of two directions.

The suggested answers for *Activity 2.16* are given below.

1. Each box in an orbital diagram represents an orbital.
2. Electron spin (spin quantum number) is represented by the direction of half-arrows. Stands only for an option to spin clockwise or anti-clockwise. 3) No. In beryllium, there are no electrons in subshells that have degenerate orbitals, so Hund's rule is not used.

The suggested answers for *Activity 2.17* are given below.

Atomic number	*Element*	*Orbital diagram*	*Full electron configuration*	*Condensed electron configuration*
11	Na		$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
12	Mg		$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
13	Al		$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
14	Si		$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
15	P		$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
16	S		$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
17	Cl		$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
18	Ar		$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$

Regarding the group assignment, you may organize the students into groups and give them as homework. *Activity 2.18* is similar to that of *Activity 2.16* except that the d subshell is involved here.

Assessment

In addition to the exercise given, ask short-answer questions at the end of each session or lesson period to assess the understanding of your students.

Students working at the minimum requirement level will be able to:

- Explain the Aufbau principle the Pauli exclusion principle and Hund's rule; and
- write ground state electronic configurations of multi-electron atoms.

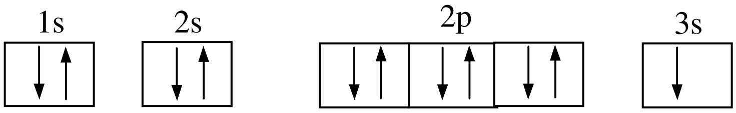
Additional Questions

* 1. How many orbitals exist in the 4f subshell?

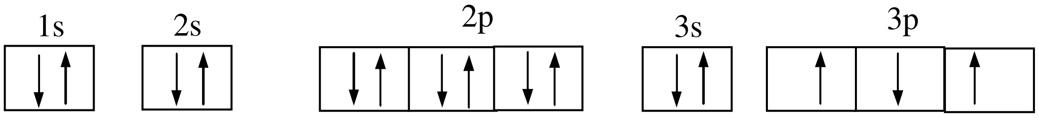
* 2. What is the lowest-numbered principal shell in which p orbitals are found?

3. Use the Pauli exclusion principle and Hund's rule to determine which of the following orbital diagrams are possible for the ground-state electron configuration and which are not. If the orbital diagram is not allowed, state why it is not.

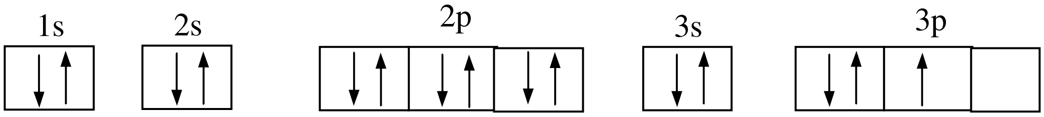
a.



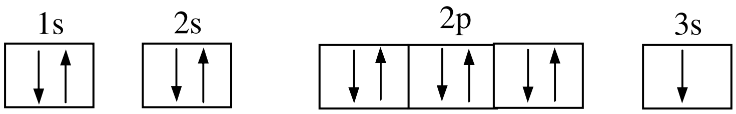
b.



c.



d.



4. None of the following electron configurations is reasonable for a ground-state atom. In each case, explain why.

   a. $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^1$ b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2\ 4p^1$

   c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2\ 4p^1$

* 5. Explain the principle (s) or rule(s) that each of the following electron configuration violates.

   a. [Ar] $2d^{10}$ b. [Kr] $4d^{10}4f^{14}5s^2$ c. [Ar] $3f^3 4s^2$

Answers to Additional Questions

1. Seven 4f orbitals

2. *n* = 2

3. a. Allowed.

   b. Not allowed. The unpaid electrons are not all the same spin.

   c. Not allowed. There should be one electron in each 3p orbital.

   d. Not allowed. The 2s electrons have the same spin.

4. a. 3s subshell fills before any electron goes into 3p subshell.

   b. 4s subshell fills before 3d subshell.

   c. 3d subshell fills completely before 4p subshell.

5. a. The subshell that fills after [Ar] is 4s. 2d does not exist.

b. The subshell that fills after [Kr] $4d^{10} 5s^2$ is 5p.

c. The subshell that fills after [Ar] $4s^2$ is 3d. 3f does not exist

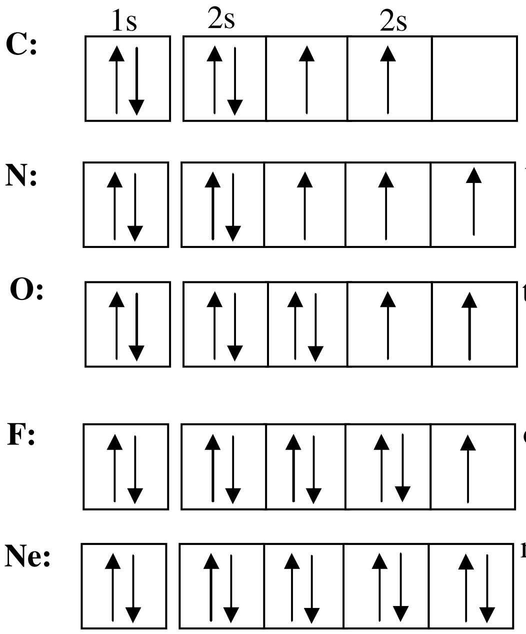
Answers to Exercise 2.8

1. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

2. $1s^2 2s^2 2p^6$

3. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

4. 1s 2s 2p



C: two unpaired electrons

N: three unpaired electrons

O: two unpaired electrons

F: one unpaired electron

Ne: no unpaired electron

Answers to Exercise 2.9

1. $Co^{3+} = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$

$Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$

$Al^+ = 1s^2 2s^2 2p^6 3s^2$

$Cr = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

$As^- = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

$Cu = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

2.8 ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE OF THE ELEMENTS

Periods allotted: 4 periods

Competencies

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

- *correlate the electron configuration of elements with the periodicity of elements;*
- *give a reasonable explanation for the shape of the periodic table;*
- *classify elements as representative, transition and inner-transition elements;*
- *explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity, and metallic character of elements within a period and group of the periodic table;*
- *write the advantages of the periodic classification of elements.*

Forward Planning

Take adequate time to study the electronic configurations and the periodic table of the elements. Make thorough preparations for explaining how the electron configuration of an element is related to the position of that element in the periodic table. Use a large-sized periodic table to indicate the location of the elements. Discuss thoroughly periodic trends like atomic size, ionization energy, and electron affinity.

Teaching Aids

The modern Periodic Table is helpful.

Subject Matter Presentation

The Modern Periodic Table

Students have prior knowledge in this topic. So, let them take enough time to be involved in deriving most of the information required and to explain the regularities observed in the properties of elements. Help the students to use the electron configuration of elements to locate their position in the periodic table and their properties. Give assignments individually and then ask them to iron out the differences between their findings.

Discuss the effective nuclear charge, screening/shielding effect and also the factors affecting ionization energy. Mention that, for anions and some neutral atoms like those of noble gases, adding an electron is an endothermic process. That is, work must be done to force an electron onto the atom. This results in the formation of an unstable anion. Use *Activities 2.19 - 2.29* appropriately in connection with your discussions.

Activity 2.19 and *Activity 2.20* are mainly concerned with the position and properties of elements in the periodic table. For instance, one can guess that metals which can form amhpoteric hydroxides are Al and Sn, and that metals which show characteristic color with flame tests are like Li, Na, K, Rb, Ca, Sr, Ba and Ra. C and Si can form covalent hydrides with the empirical formulas CH_2 and SiH_2. He, Ne, and Ar (with their respective boiling points) do not form oxides, hydrides, or chlorides. N and P form oxides of empirical formulas N_2O_5 and P_2O_5. On hydrolysis, the oxides of N yield a much more acidic solution than that of P. Silicon forms an oxide which has high melting point and is an important raw material of glass. Give the students a chance to guess the name and location of the element in the periodic table.

Elements in a group form a family with some variations in physical properties. The chemical properties of an element depend on its electrons. Since elements in a group have the same number of electrons, their chemical properties are the same.

To keep the periodic table at maximum width of 18 members, the f-block elements are placed below the main body of the periodic table. Isotopes of an element are located in the same period and group. The properties of compounds formed can be predicted in their location in the periodic table.

Classification of the Elements

Introduce the classification of elements based on valence electron(s) entering the orbital of lower energy. It is advisable to use both the Roman numeral; A and B Group – I A to VIIIA for representative(s and p block elements), I B to VIII B for transition (d – block) elements as well as the latest method of classification as Group 1–18.

You may use *Activity 2.21* to help students actively participate in discussion. The elements in the periodic table are classified based on the orbitals being filled in the Aufbau process, like s, p, d, and f. They are all separate and distinct. There is no need to learn all the elements in the periodic table. Some industrially applicable elements and their compounds can be studied based on their periodic properties.

Periodic Properties

Activity 2.22 may be used to help motivate students' participation in discussion. The following suggested answers are given below.

1. The size of elements in general tends to decrease across a period (from left to right) because the effective nuclear charge increases from left to right in a period of representative elements in the periodic table.

2. Because metal atoms have small number of electrons in their valence shells. Except for hydrogen and helium, all s-block elements are metals.
3. Because all their valence electrons are paired and fulfilled an octet. Therefore, these elements are least reactive among all the elements in the periodic table.
4. A shield is in conversational English a broad piece of metal, wood or hide used as a protection against being hit by a thrown spear. This can be used as a sort of analogy to help students easily understand shielding effect/screening effect. Electrons (valence specially) which are the subjects of losing and gaining regions experience an attraction to a net-positive charge of the nucleus. So, the core electron(s) are shield (protect) the valence electron(s) from the intensive attraction of the nucleus (the spear in the analogy given). This protection by the core electrons is known as shielding effect. The "fighting" is now between the positive charge of the nucleus and the core electrons which are not numerically equal. The positive charge always exceeds the core electrons, and the resultant is known as the effective nuclear charge.

After harmonizing students' discussion on *Activity 2.22*, you may use *Activity 2.23* for further discussion. The following suggested answers are given below.

1. Screening increases with increasing atomic number.
2. In a series of transition elements, electrons join the core electrons, not valence shell electrons. As a result, the effective nuclear charge for Cr and, Mn remains more or less constant instead of increasing.
3. On the left-most corner of the periodic table, with least effective nuclear charge.
4. Though isoelectonic ions have the same electrons, their nuclear charges are different. So we do not expect them to have the same radii.
5. Because electrons could be found at infinite distances from the nucleus.
6. a. decreases
   b. Increases for the same reasons explained above
7. a. The nuclear charge of He is greater than H. They have similar effective nuclear charges
   b. The sizes of atoms increase down a group. He is below He in Group 8

Activity 2.24 may be used to facilitate discussion on ionization energy. The suggested answers are as follows.

1. Because IE_3 is when aluminum loses its three valence electrons and attains the Ne configuration-stable. IE_4 energy required to remove an electron from a stable (octet) configuration. Therefore, a big difference is expected.
2. Both S and P are in Period 3 of the periodic table. Sulfur removes one electron from the paired p subshell, while phosphorus removes one electron from the half-filled p subshell.
3. After lithium removes one electron from its outermost shell, it attains the configuration of He. Be must remove two electrons to have an electron configuration like that of He.
4. The difference between IE_2 and IE_3 is greater because IE_3 is the energy required to remove an electron from a filled subshell.

Students may apply cooperative learning methodology in the group assignment suggested. Have the students form groups. Help students to work on the group assignment as homework.

Activity 2.25 may be used to facilitate discussion on electron affinity. The suggested answers are as follows.

1. In the case of Ne, an electron will be added onto a completely filled subshell. Therefore work must be done to add an electron onto the last subshell. Electron affinity of F is negative because fluorine must lower its energy by gaining one electron to have an electron configuration like that of the noble gas Ne.
2. Group 7
3. This is due to the electron configuration of Si v_s P. An electron added to P must go into a p orbital that already contains an electron. The electron that is already there repels the added electron, and less energy is released during this electron addition, so the electron affinity is lower.

Activity 2.26 may be used to facilitate discussion on the concept of the second electron affinity. The suggested answer is as follows.

It is not hard to see why EA_2 is a positive quantity. Here an electron approaches an ion with a net charge of -1. It is strongly repelled, and work must be done to force the extra electron on to the $O^-(g)$ ion. In fact, we only except O^{2-} ion to form in situations where the large energy expense to create it can be offset by another energetically favorable process.

Activity 2.27 may be used to facilitate discussion on the concept of electronegativity. The suggested answers given below may be used to harmonize the discussion.

1. Electronegativity is the power of attracting shared electrons between two atoms, whereas electron affinity is the energy released when an electron is added onto a gaseous atom or ion.
2. Electronegativity decreases down a group in a family of a periodic table as size increases.

Activity 2.28 may be used to facilitate discussion on the concept of metallic character. The suggested answers given below may be used to harmonize the discussion.

Metallic character increases down a group and deceases across a period. Atomic size increases down a group and decreases across a period. Ionization energy decreases down a group and increases across a period. One can see from these periodic properties that metallic character and atomic size have the same trend.

Activity 2.29 may be used to facilitate discussion on the concept of advantage of periodic classification. The suggested answers given below may be used to harmonize the discussion.

a. Leftmost side of the periodic table.

b. In group (right side of the periodic table).

c. Group 6, period 4.

Assessment

In addition to the exercise given, ask short-answer questions at the end of each session or lesson period to assess the understanding of your students.

Students working at the minimum requirement level will be able to:

- Classify elements as representative, transition and inner-transition elements;
- Explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity and metallic character of elements within a period and group of the periodic table; and
- Write the advantages of the periodic classification of elements.

Additional Questions

1. Which of the following hydrogen orbitals have identical energies?

   a. 1s, 2s, 2p b. 3s, 3p, 3d c. $3p_x$, $3p_y$ $3p_z$

* 2. What is the configuration of the electrons in the outermost shell in the Group 3A elements? In Group 5A elements?

* 3. What subshell(s) is (are) being filled in each of the following regions of the periodic table?

a. Group 1A and 2A b. Group 3A through 7A

c. The transition elements d. The lanthanides and actinides

* 4. Give the periodic table period number and group number for the element whose atoms have these electron configurations:

a. $1s^2 2s^2 2p^6$ b. $1s^2 2s^2 2p^6 3s^2 3p^6$

c $1s^2 2s^2 2p^6 3s^1$ d. $1s^2 2s^2$

e. $1s^2 2s^2 2p^6 3s^2 3p^1$

5. What is the maximum number of electrons that go into a d orbital? A d subshell?

6. Write out the spdf notation for the electrons in the highest principal energy level for each of the following elements.

a. Cs b Se c. In

* 7. Write out the complete spdf notation for each of the following elements. Identify the core electrons and the valence electrons in each.

a. Si b. Rb c. Br

* 8. Arrange each of the following pairs of elements in order of increasing atomic radius.

a. Cl or S b. Al or Mg c As or Ge d. Ca or K

9. Arrange each of the following sets of elements in order of increasing first ionization energy.

a. Ca, Mg, Sr b. Cl, P, S c. As, Ge, Sn d. Br, Cl, Se

10. Arrange each set of elements in order of increasing atomic radius and explain the basis for this order.

a. Al, C, Si b. Ca, Rb, Sr

*11. Propose a probable electron configuration for the unknown and as yet undiscovered element with Z =114.

12. The production of gaseous chloride ions from chlorine molecules can be considered to be a two-step process in which the first step is $Cl_2 (g) \rightarrow 2Cl (g)$ $\Delta H = +242.8 kJ$

What is the second step? Is the overall process endothermic or exothermic?

*13. In the students' textbook, ionization energies are given in the unit kJ/mol. Another way of expressing these quantities is in terms of a single atom rather than a mole of atoms, through the unit electron volts per atom, eV/atom. Use physical constants and other data from any source to show that 1 eV/atom = 96.49 kJ/mol

Answers to Additional Questions

1. If an electron of 1s hydrogen is excited to 3p, and all the 3p three orbitals are degenerate. Therefore choice b and c are possible answers.

2. Group 3A: ns^2np^1

   Group 5A: $ns^2np_x^1np^1_ynp_z^1$

1. a. ns b. np c. (n – 1)d d. (n – 2)f

2. a. period 2, Group 8A b. period 3, Group 8A

   c. period 3, Group 1A d. period 2 Group 2

   e. period 3 Group 3A

3. d orbital, two electrons; d subshell, ten electrons

4. a. $6s^1$ b. $4s^24p^4$ c. $5s^25P^1$

5. a. $1s^2\ 2s^2\ 2p^6$, core; $3s^23p^2$, valence

   b. $1s^2\ 2s^2\ 2p^6\ 3s^23p^63d^{10}4s^24p^6$, core; 5s; valence

   c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$, core; $4s^24p^5$, valence

6. a. Cl < S b. Al < Mg c. As < Ge d. Ca < K

7. a. Sr < Ca < Mg b. S < P < Cl c Sn < Ge < As d Se < Br < Cl

8. a. C < Si < Al atomic radius increases from top to bottom in a group of elements, and from right to left in a period of elements.

   b. Ca < Sr < Rb (for the same reason as in part (a)).

9. $[Rn]\ 5f^{14}6d^{10}7s^2\ 7p^2$

10. $Cl(g) + 1e^- \rightarrow Cl^-(g)$; $\Delta H = EA = -349$ kJ. For the overall process,

    $\Delta H = +242 - 349\ kJ = -106$ kJ. The overall process is exothermic.

11. $$\frac{1eV}{atom} \times \frac{6.02 \times 10^{23}\ atoms}{mol} \times \frac{1.6{\times}10^{-19}C}{electron} \times \frac{1kJ}{10^3J} = 96.49kJ/mol$$

Answers to Exercise 2.10

1. F < S < P < As
2. Na < Mg^+ < Al^{2+} < Si^{3+}

Answers to Exercise 2.11

1. Generally it is possible to explain the irregularities in the trends across a period. Note, for example, that aluminum, with Z = 13 and the electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^23p^1$ has a smaller first ionization energy than does magnesium (Z= 12), which has the electron configuration $1s^22s^2\ 2p^6\ 3s^2$. The 3P electron of Al is at a

higher energy than the 3s electrons of Mg and is therefore easier to remove. This kind of discontinuity occurs generally in proceeding from a Group 2A element to Group 3A elements and from 5A to 6A elements.

2. In the solid and liquid state of elements energy applied requires to overcome the lattice and vaporization energies before removing electron from an atom.

3. Potassium (Z=19) and electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^1$ removes its 4s electron with great ease, attaining the noble gas configuration (Ar), and it is extremely difficult to remove the second electron from a stable configuration.

   Calcium (Z=20) and electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ removes the two electrons to become like that of noble gases in its configuration. This is the reason why the energy difference between 1st and 2nd is small as compared to potassium.

4. The valence electron of K requires lower energy to be removed since K is greater in size than Li.

5. Ga < Ge, Sb < Br, K < Cr, Sr < Mg, O < N.

Answers to Review Questions and Problems

1.	None	2.	A	3.	C	4.	B	5.	D
6.	B	7.	D	8.	A	9.	D	10.	B
11.	C	12.	B	13.	A	14.	B	15.	D

16. a. protons b. electron c. electron
    d. neutron e. electron

17.

Element	*Mass №*	*Protons*	*Neutrons*	*Electrons*
Neon	20	10	10	10
Nitrogen	14	7	7	7
Gallium	70	31	39	31
Nickel	59	28	31	28
Iron	56	26	30	26

18. a. $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$

b. Atoms of the same element with different neutrons but the same number of protons.

c. 63.55 m_u

19. a. Same number of protons but different numbers of neutrons- isotopes.

b. The chemical properties of an element is determined by its electrons. The two particles have the same number of electrons. Therefore, they have similar chemical properties

20. Na < Li < O < N < Ne

21. a. The energy required to remove an electron from a p orbital is less than that of removing an electron from an s orbital.

b. The greater repulsion experienced by the fourth 2p electron makes it easier to remove – pairing starts at oxygen.

c. Because lithium is smaller in size than sodium.

22. a. Vanadium b. $[Ar]4s^2 3d^3$

23. Al < Mg < Ca < Ba < Cs

24. Potassium (Z= 19) and electron configuration $[Kr]\ 4s^1$ removes its 4s electron making it like that of the nearest noble gas, whereas calcium (Z=20) and electron configuration $[Kr]4s^2$ removes the two 4s electrons, giving it an electron configuration like that of noble gas.

25. a. $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-}$

b. $Na^+ < Ne < F^- < N^{3-} < C^{4-}$

c. F < C < B < Be < Li

d. $Ca^{2+} < K^+ < Cl^- < S^{2-} < As^{3-}$

26. a. 5.09×10^{14} Hz b. 3.37×10^{-19}J

27. a. 388.9 nm b. 7.71×10^{14} Hz

28. $E_n = -R_H \times \frac{1}{n^2}$; -2.42×10^{-19} J

29. 1.03×10^{-7} m

30. $\lambda = \frac{h}{mv}$; 2.64×10^{-10} m

31. a. 3p b. 4s c. 3d d. 5f

32. a. $[Ar]3d^5$ b. $[Ar]4s^2 3d^3$ c. $[Ar]3d^3$ d. $[Ne]3s^2$

33. Zr and Fe

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UNIT 3

CHEMICAL BONDING AND STRUCTURE

UNIT OVERVIEW Total Period allotted: 32 periods

This unit begins with section 3.1, which is an overview of how atomic properties give rise to the three main ways that atoms bond. Section 3.2 deals with ionic bonding and ionic compounds as well as the electron configuration of the ions. A general method of writing Lewis electron dot formulas for ionic bonds is discussed in this section. Also, the section further discusses the factors affecting ionic compounds. Section 3.3 looks at covalent bonding and covalent compounds, followed by representations of covalent bond (drawing Lewis structure).

In the same section, subsection 3.3.1 deals with molecular geometry. The focus of this subsection is the shapes of molecules and its effect on the properties of a molecule. It starts by describing the valence shell electron pair repulsion model (VSEPR), which allows converting two dimensional Lewis structures into three-dimensional geometries. Subsection 3.3.2 will turn your attention to the attractive and repulsive forces among the particles – molecules, atoms or ions in a sample of matter. It is an intermolecular force.

Section 3.4 deals with the bonding in metals which occurs among large numbers of metal atoms. In this section, you will discuss a qualitative model. A quantitative aspect of it is considered to be not important here. Two models (electrons-sea model and molecular orbital or band models) are used to explain the bonds in metals.

Chemical bonding models (valence bond and molecular orbital) for covalent compounds are discussed in section 3.5. Finally in section 3.6, the four important classes of crystalline solids with related properties are presented.

Unit Outcomes

After completing this unit, students will be able to:

- *understand that a chemical bond is an attractive force between particles;*
- *demonstrate an understanding of the formation and general properties of substances containing ionic, covalent and metallic bonds;*
- *draw Lewis structures for simple ionic and covalent compounds;*
- *understand the origin of polarity within molecules;*
- *describe the formation and nature of hydrogen bonds, dipole-dipole forces and London forces;*

- *know the three different but related bonding models (Lewis model, valence bond model and molecular orbital model) and recognize the usefulness of the bonding theories in explaining and predicting molecular properties (bond angle, bond length, bond energy, etc);*
- *explain how the properties of a substance (solid or liquid) depends on the nature of the particles present and the type of intermolecular forces;*
- *appreciate the importance of intermolecular forces in plant and animal life;*
- *explain how the valence shell electron pair repulsion (VSEPR) model can be used to predict molecular shape;*
- *know the types of crystalline solid (ionic, molecular, covalent network, or metallic) formed by a substance and describe their properties;*
- *conduct experiments to observe and analyze the physical properties of different substances to determine the type of bonding present; and*
- *apply the scientific enquiry skills for this unit: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, asking questions, applying concepts, relating cause and effect and making generalizations.*

Main Contents

3.1 Introduction

3.2 Ionic Bonding

3.3 Covalent Bonding

3.4 Metallic Bonding

3.5 Chemical Bonding Theories

3.6 Types of Crystals

3.1 INTRODUCTION

Periods Allotted: 1 period

Competencies

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

- *describe the reason why atoms form chemical bonds;*
- *state octet rule;*
- *define chemical bonding;*
- *describe the types of chemical bonding and their mechanisms of the bonding processes.*

Forward Planning

Read the material and relate your plan with what the students learned in lower grades.

Teaching Aids

No additional teaching material is suggested.

Subject Matter Presentation

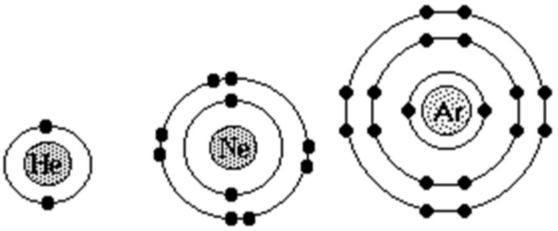
First and foremost, it is very important to engage students in *Activity 3.1*, which is intended to remind them their knowledge of chemical bonding.

The suggested answers for *Activity 3.1* are given below.

1. The number of electrons in the outermost shell or orbit of an atom is ready to combine with other atoms to form a bond. Except in transition metals, all other elements in the periodic table should have electrons in the first orbit as 2 and in all other outer shells as 8 electrons
2. By combining in specific compounds, the total energy of the atoms involved is reduced. The greater this reduction in energy (between the free isolated atoms and the combined atoms in molecules), the more likely the atoms combined will become stable.
3. The forces that hold atoms together in molecules and that keep ions in place in solid ionic compounds is called a chemical bond
4. In a molecule there will always be integer numbers of a certain atom, for example you can never have 3.5 oxygen atoms or 2.493 chlorine atoms in a compound. Atoms are singular units of molecules. Therefore a ratio of integers is always present for all types of atoms in a molecule.
5. For covalent bonds, bond energies and bond lengths depend on many factors: electron affinities, sizes of atoms involved in the bond, differences in their electronegativity, and the overall structure of the molecule. There is a general trend, in that the shorter the bond length, the higher the bond energy. However, there is no formula to show this relationship, because the variation is widespread. From a table of values, we cannot grasp the trend easily. The best method for seeing the trend is to plot the data on a graph. In a discussion of bond energies, this link has shown how energy varies as two H atoms approach each other in the formation of an H-H covalent bond:

Activity 3.2 may be used to introduce the concept of octet rule. The following suggested answers are given below.

1. Some atoms are very reluctant to combine with other atoms and exist in the air around us as single atoms. All other atoms, therefore, bond together to become electronically more stable that is, to become like noble gases in electron arrangement. Bonding produces new substances and usually involves only the 'outer shell' or 'valence' electrons. Atoms can bond in two ways.
2. What is common amongst these atoms is the making of eight electrons (octet) on their outermost shell (except He). They are all under what we now call Group 18 (VIIIA).
3. They have very stable electron arrangements e.g., 2, 2, 8 and 2, 8, 8-because their outer shells are full. The first three are shown in the diagrams below, which indicates why noble gases are so reluctant to form compounds with other elements.



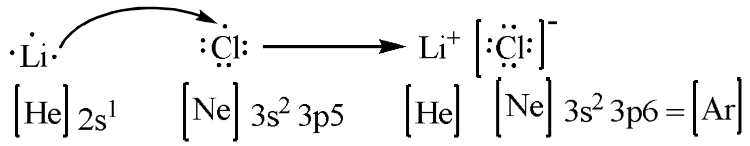
helium (2) 2 neon (10) 2.8 Ar (18) 2, 8, 8

(Atomic number) electron arrangement

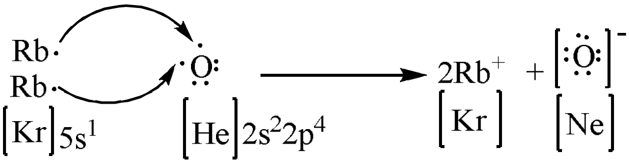
4. These elements were known as the inert gases (now noble gases); no chemical reactions were known for any of them. But, these days, quite a large number of Xe compounds are known. Evidence exists of krypton and radon compounds as well. But no compounds have yet been made of helium, neon and argon.
5. The driving force for all types of chemical activity is energetic in character-lowering of energy is the sign of stability. You can give a sort of analogy here like: a ball rolling down a roof will come to a new position of rest at a lower energy. So a particular arrangement of atoms undergoes spontaneous chemical reaction to come to a new arrangement with a lower energy.
6. Atoms will react to get in to the most stable state possible. A complete octet is very stable because all orbitals are full. Atoms with greater stability have less energy, so a reaction that increases the stability of the atoms will release energy in the form of heat or light. Reactions that decrease stability must absorb energy, getting colder.

Activity 3.3 is designed to introduce the concept of chemical bonding. You may use the following suggested answers to harmonize the discussion.

1.

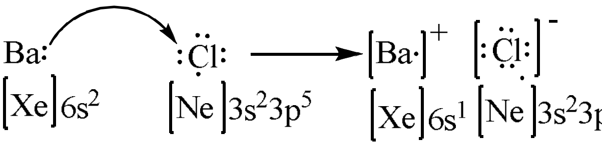


2. a.



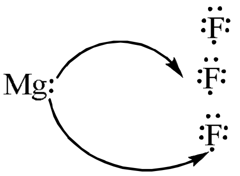
It is a likely formula because it satisfies the octet rule.

b.



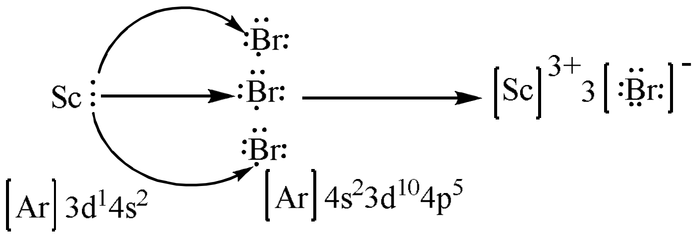
It is not a likely formula because Ba does not exhibit 8 valence electrons

c.



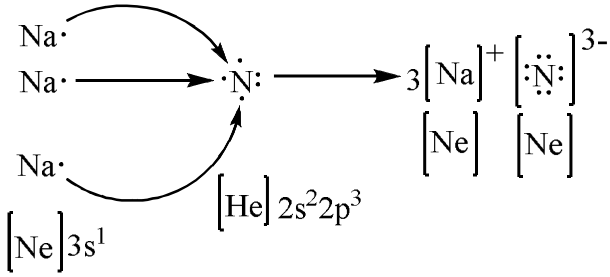
It is unlikely formula because one Mg can satisfy only two F atoms.

d.



It is a likely formula.

e.



It is a likely formula.

Assessment

You may ask question of your own at the end of the lesson. Use the additional exercises as required.

Additional Questions

1. How does an atom become an ion?

* 2. Two or more atoms combine to form

a. An anion   b. A molecule

c. An ion   d. A cation

* 3. Inert gases do not react with any other element because

a. The temperature is not high enough

b. Their outermost electron level is filled with 8 electrons

c. Their nucleus is very small

d. The systems pressure is not enough

* 4. Atoms take part in bond formation to

a. Attain a stable electronic configuration

b. Increase their charge density

c. Increase their energy

d. Neutralize their charge

* 5. How many electrons are required in the outermost shell of an element for it to be stable?

a. 8   b. 6   c. 5   d. 4

* 6. Which of the following compounds is ionic?

a. CH_4 b. H_2O c. HCl   d. NaCl

Choose from the following terms and complete the sentences below. Some of the terms may not be used.

nucleus
outer
atomic structure
first
second
electron dot
proton
electron
proton

* 7. An element can have from one to eight electrons in its ___________energy level.

* 8. The closer a(n) _______ is to the nucleus, the stronger the attractive force.

* 9. An atom's ___________ contains its protons and neutrons.

* 10. The ________________ energy level of an atom can hold only two electrons.

Answers to Additional Questions

1. If an atom loses or gains electrons, it is no longer called an atom.
   It is called anion. An ion has a full outer shell of electrons. An ion has an electric charge.

2. b
3. b
4. a
5. A
6. d
7. Second
8. electron
9. nucleus
10. first

3.2 IONIC BONDING

Period Allotted: 5 periods

Competencies

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

- *define ionic bonding;*
- *use Lewis electron dot symbols to depict main group elements;*
- *describe ionic bonding using Lewis electron dot symbols;*
- *list the favorable conditions for the formation of ionic bond;*
- *explain the formation of ionic bonding;*
- *give examples of ionic compounds;*
- *define Lattice energy;*
- *calculate lattice energy of ionic crystal from a given data using the Born-Haber cycle;*
- *discuss the exceptions to octect rule;*
- *describe the properties of ionic bonding;*
- *carry an activity to demonstrate the effect of electricity on ionic compounds (PbI_2 and NaCl);*
- *carry an activity to investigate the melting point and solubility of some ionic compounds (NaCl and $CuCl_2$).*

Forward Planning

Considering why bonds are formed between atoms. A necessary condition for the formation of an ionic bond is the transfer of electrons. Plan how to present this topic so that your students can grasp the idea and see the concepts and examples of chemical bonding –ionic bonding for the time being.

Apply what you think to be impressive and illustrative concepts and examples to teach factors affecting the formation of ionic compounds, properties of ionic compounds, the Born-Haber cycle, which deals with energetic, and how to write formulas for ionic compounds.

Teaching Aids

You may use flash cards to highlight the key words used in this section. Diagrams sketched on flip charts will help you also.

Subject Matter Presentation

Arrange students in groups and help them to participate actively in group discussions on *Activity 3.4*. Once again, this activity is among those which are intended to remind students of their chemistry lessons in Grade 9. You may use the following hints to harmonize the discussion.

The elements involved in ionic bonding are metals and nonmetals. Ionization energy is defined as the energy required to remove an electron from a gaseous atom or ion to a distance of infinity. Electron affinity proceeds with evolution of heat, which plays a role in making the ion more stable than the separated atom in the formation of an ionic bond.

Two ionic bonds will result from the formation of $MgCl_2$ from magnesium and chlorine.

Lewis Electron Dot Symbols

Discuss with the students Lewis symbols for electrons. Take into consideration their knowledge of grade nine concepts that Lewis symbols are generally written only for the main group or representative elements. Ask them why this is easy to do. You are aware that the number of valence electrons for these elements is equal to the group number. Investigate the knowledge of students do they have the idea of how electron spin and other fine points about electronic that structure had been proposed when Lewis formulated his theory? You can use *Activity 3.5* number 2 for better illustration. Stress the concept that, in a number of cases, Lewis symbols lead to a better prediction of chemical bonding than do ground-state electron configurations.

You may use the following suggested answers to harmonize the discussion.

1. a. $[Ar]$, $[K]^+$

b. $[\mathrm{Ar}],\left[:\ddot{\underset{\cdot\cdot}{\mathrm{S}}}:\right]^{2-}$

c. $[\mathrm{Ne}],\left[:\ddot{\underset{\cdot\cdot}{\mathrm{F}}}:\right]^{-}$

2. Notice that, for example, for elements in Groups 2A, 3A, and 4A, the Lewis symbols do not affect the electron pairing we would expect in ground-state electron configurations that is, that Be actually has no unpaired electrons, B has one, and carbon has two. But in their Lewis electron-dot symbols, Be has two, B has three, and C has four unpaired electrons which are available for pairing purposes.

Formation of Ionic Bonding

Once again, do not forget to bear in mind that the students have some background knowledge of ionic bond formations and start your discussion by asking questions like: What is ionic bonding? How does it differ from other types of bonding? What are the factors affecting ionic bond formations? You may ask the students to use sketches or diagrams to show what they visualized. You are also advised to use the additional exercises in this teachers' guide to assess their mastery of the concept.

Exercise 3.3 and the suggested activity (*Activity 3.6*) can be used for further stabilization of the subject matter. The purpose of this activity is to motivate students for the search of the coulombic attractive force and its relations with charge and size of an ion.

Discuss the energy changes in the formation of ionic bonds. Emphasize the idea that there are three steps involved in the formation of an electrovalent (or ionic) bond. These are

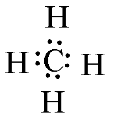
- removal of electron(s) from one atom. In this step, energy equal to the ionization energy is absorbed.
- gaining of electron(s) by the other atom. In this step, energy equal to the electron affinity is released.
- combination of cations and anions. Coulombic forces of attraction hold these ions together. In this step, energy equal to the lattice energy is released.

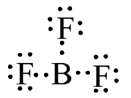
Activity 3.6 is prepared to help students grasp the concepts of factors affecting formation of ionic bonding. You may use the following suggested answers to harmonize the discussion.

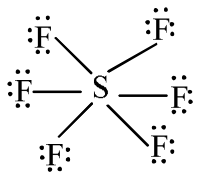
1. The lattice energy of an ionic compound is directly proportional to the charges involved and inversely proportional to the square of the radius between the two ions.
    a. LiCl because Li is by much smaller than Cs in size.
    b. MgO the charges on Mg and O are twice that of Na and F. Charge outshined size.
    c. BaS for similar reasons as discussed in (b).

Activity 3.7 refreshes students' prior knowledge of the octet rule, exceptions to the octet rule, and the Lewis structure. Let the students discuss on it and come to their conclusions.

During their discussion of *Activity 3.7* number 1, if they conclude that Fe^{3+}, Cu^{2+}, and Li^{+} are the cations that do not acquire noble gas octet configuration, they are quite right. Then ask them to draw the Lewis structure for the suggested molecules (number 2) though these species are not ionic. This activity is merely to evaluate their ability to draw a Lewis structure, recalling the rules from their Grade 9 chemistry lesson.







Soon after this activity is discussed in groups and a sort of consensus is reached between different groups, start to discuss on exceptions to the octet rule in ionic compounds. One way or the other, this idea is incorporated during very few lesson periods. Discuss in particular those which exhibit less than octet and above octet electron accommodations in their valence shells.

Properties of Ionic Compounds

Let the students read the notes during the lesson; give them guiding questions, for example, do ionic compounds conduct electricity, or do they dissolve in water? to motivate their discussion and arguments during reading. Four experiments are suggested. First try to perform the experiments yourself ahead of time so that it will be easy when you conduct them with the students. Arrange appropriate time for the experiments maybe one at a time or two, depending on the situations in your school. You may use the suggested exercises as class work or homework, as required. You are also advised to prepare additional questions for students so you can assess their mastery of the concepts contained in the content.

Experiment 3.1is designed to investigate the solubility of NaCl and $CuCl_2$. The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

When an ionic crystal such as NaCl is placed in water, a dissolution process will occur. Initially, the positive and negative ions are only attracted to each other. The water molecules are hydrogen bonded to each other. If the crystal is to dissolve, these bonds must be broken. Negative chloride ions on the surface are attracted by neighboring positive sodium ions and by the partially positive hydrogen atom in the polar water molecule. Similarly, the positive sodium ions are attracted by both chloride ions and the partially negative oxygen atom in the polar water molecule.

How does the polarity of water allow it to effectively dissolve ionic compounds?

To dissolve an ionic compound, the liquid must be able to effectively stabilize the ions that result from breaking the ionic bond.

Solvent	***NaCl***	***$CuCl_2$***
Water	Soluble	Soluble
Ethanol	Insoluble	Insoluble
Hexane	Insoluble	Insoluble
Benzene	Insoluble	Insoluble

Experiment 3.2 is designed to study the effect of heat on ionic compounds. The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

The melting point of NaCl is 801^0C. This temperature is very high and impossible to obtain in the lab. Similarly, $CuCl_2$ needs high temperature to melt; as a result, it is generally difficult to determine the melting points of these two and other high melting point ionic compounds.

Experiment 3.3 is designed to test the electrical conductivity of molten ionic compounds. Guide the students arranging the set up in groups, as shown in the students' textbook (Figure 3.4). The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

The bulb doesn't glow when lead (II) iodide crystals are used as an electrolyte. This indicates that solid ionic compounds do not conduct electricity. But when they are molten, the bulb glows, indicating that they are conductors since ions which are responsible for the conduction of electricity are produced during melting.

Ionic compounds (which are composed of elements of low ionization energy and high electron affinity-metals and non-metals) conduct electricity in molten state. The type of bonding which exists in these compounds is known as ionic bonding.

The purpose of Experiment 3.4 is to test the electrical conductivity of aqueous solution of some ionic compounds. Guide the students arranging the set up in groups, as shown in the students' textbook (Figure 3.5). The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

Solutions of ionic compounds conduct electricity because they also dissociate into cations and anions. An electrochemical reaction takes place in the presence of a power source, and these reactions at the respective electrodes may increase the mass of the electrode due to discharge of the cation and deposition of the metal on the cathode.

When molten ionic compounds are electrolyzed, cations usually discharge at the cathode and anions at the anode. Note that melting produces positive and negative ions-there is no preference. In the case of solutions of electrolytes, three or more ions can be produced, and only one ion is preferentially discharged at the respective electrode. Take, for example sodium chloride.

$$\mathrm{NaCl + H_2O} \longrightarrow \underbrace{\mathrm{Na^+ + H^+}}_{\substack{\downarrow \\ \text{to cathode}}} + \underbrace{\mathrm{Cl^- + OH^-}}_{\substack{\downarrow \\ \text{to anode}}}$$

$$\mathrm{NaCl} \xrightarrow{\text{heat}} \underset{\substack{\downarrow \\ \text{to cathode}}}{\mathrm{Na^+}} + \underset{\substack{\downarrow \\ \text{to anode}}}{\mathrm{Cl^-}}$$

Assessment

You are required to assess students' active involvement in each of the activities. At the end of each lesson, ask short-answer questions, as you implement your lesson plan. Your student performance record list is helpful in all your assessments. Don't forget to appreciate students working above the minimum requirement level. You are required to assist those working below the minimum requirement level.

Continuous assessment during the teaching-learning process is very important. Moreover, different varieties of questions, when necessary, can be asked. For instance, the electron transfer during the formation of ionic bonding, the Lewis formula for reactions involving the transfer of electrons, violations of the octet rule, contemplating and calculating energy forms involved during ionic bond formation and the like are subjects for questions which can be asked.

Additional Questions

1. How is an ionic bond formed?

*2. What is the chemical formula of sodium chloride?

*3. Deduce the electronic configuration and formula for the ionic compound formed by the elements Al and F; draw its Lewis structure (Al = 13, F = 9).

4. What is the difference between an ionic and a molecular compound?

5. What is a crystal lattice in an ionic compound?

*6. What is the chemical formula of lead chloride?

*7. What is the formula for potassium sulfate? Is the compound ionic or covalent?

8. What is the ionic compound name for lithium and argon?

9. How can electronegativity be used to distinguish between an ionic bond and a covalent bond?

Answers to Additional Questions

1. Ionic bonding occurs between a pair of atoms when one of the atoms gives up its valence electrons to the other.

2. NaCl

3. The electronic configurations of aluminum (At. no. 13) and fluorine (At. no. 9) atoms are 2,8,3 and 2,7, respectively. Aluminum has 3 electrons in excess of its nearest noble gas (neon) configuration. Fluorine is one electron short of its nearest noble gas (neon) configuration. Therefore, one atom of aluminum will transfer its three electrons to three atoms of fluorine (one of each) as shown below:

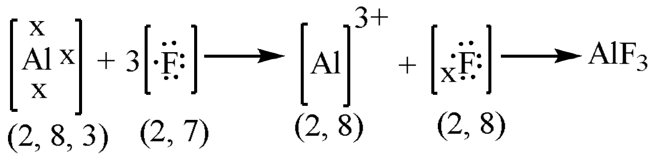
$$\underset{(2.\ 8.\ 3)}{\mathrm{Al}} \longrightarrow \underset{(2,\ 8)}{\mathrm{Al}^{3+}} + 3e^-$$

$$\underset{(2,\ 7)}{3\mathrm{F}} + 3e^- \longrightarrow \underset{(2,\ 8)}{3\mathrm{F}^-}$$

The Al^{3+} and three F^- ions then form ionic bonds between them,

$$\mathrm{Al}^{3+} + 3\mathrm{F}^- \longrightarrow \mathrm{AlF}_3$$

In terms of Lewis dot structure,



4. Ionic compounds are between metals and nonmetals. Molecular (covalent) compounds are only between nonmetals.
5. A regular and repeated three-dimensional pattern of arrangement of positive and negative ions.
6. The chemical formula of lead chloride is $PbCl_2$.
7. The formula of potassium sulfate is K_2SO_4, ionic.
8. Argon is an inert gas, and an inert gas does not want to combine with anything else.
9. The larger the difference in electronegativity, the more ionic properties a bond is said to have.

Answers to Exercise 3.1

1. a. Potassium, by losing one electron, and chlorine, by gaining the lost electron, form an ionic bond.
   b. Magnesium losing two electrons; oxygen gaining two.
   c. Two sodium atoms losing two electrons; one oxygen atom gaining the two lost electrons.
2. Only calcium because chlorine is a non-metal, and calcium is a metal. An ionic bond is between a metal and a non-metal. Carbon, oxygen and silicon are non-metals.
3. The number of electrons lost or gained by an atom of any element is called its electrovalency. The elements that give up electrons to form positive ions are said to have positive valency, while the elements which accept electrons to form negative ions are said to exhibit negative valency. What is going on in ionic bonding is this losing and gaining process. Therefore, equivalently this process is known as electrovalent bonding.
4. Three types of chemical bonding – ionic, covalent, and metallic.
5. The net attractive force that holds positive and negative ions together in a crystalline solid is called an ionic bond. It is formed by transfer of electrons from one atom to another atom. Electrons are not actually gained or lost by the atoms involved in covalent bonding. They pass into a “shared” state. A metallic bond is one in which all of the metal atoms share with each other a cloud of electrons. The electrons that make up that cloud originate from the outermost energy levels of the atoms.
6. Ionic compounds are usually crystalline solids with high melting and boiling points. They are conductors of electricity when molten or dissolved. They are soluble in inorganic solvents (water is the most common solvent for ionic

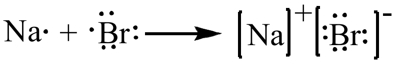
compounds) but insoluble in organic solvents like benzene, ethanol and carbon tetrachloride.

7. Recall that network covalent substances have some of the highest melting and boiling points known. Since melting a network solid requires the breaking of covalent bonds, we can infer that covalent bonds are the strongest of the forces of attraction. Ionic compounds tend to have melting points well above 300°C. They are all solids in pure form at room temperature. A rule of thumb is that ionic bonds are about half as strong as covalent bonds.

Metals have a wide range of melting and boiling points. For example, mercury is a liquid at temperatures above -39°C. Tungsten, on the other hand, does not melt until the temperature reaches 3410°C. The wide range of properties does not allow us to make general statements about the relative strength of metallic bonds versus other chemical bonds.

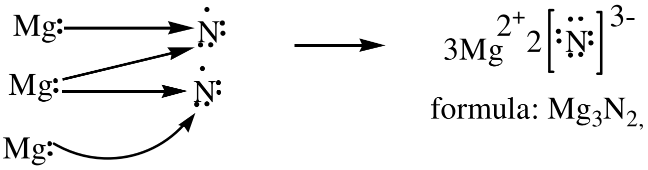
Answers to Exercise 3.2

1.



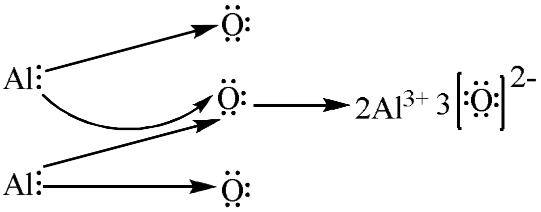
Formula: NaBr

2.



formula: Mg_3N_2,

3.



Answers to Exercise 3.3

1. a. $SrBr_2$ (strontium bormide) b. CaS (calcium sulphide)
   c. LiF (lithium fluoride)

2. a. Group 2 b. Group 6 c. Group 3 d. Group 6

3. $\Delta H_f^o = -1047 kJ/mol + \left(161 + 520 + \frac{154}{2} - 328\right) kJ/mol = -617\ kJ$

4. – 669 kJ

5. $-795.8 - [178.2 + 243.4 + 590 + 1145 + (-348.7 \times 2)] = -2255\ kJ$

3.3 COVALENT BONDING AND MOLECULAR GEOMETRY

Period Allotted: 15 periods

Competencies

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

- *define covalent bonding;*
- *explain the formation of covalent bonding;*
- *give examples of covalent molecules;*
- *draw Lewis structures or electron dot formulas of some covalent molecules;*
- *illustrate the formation of coordinate covalent bonding using examples;*
- *define resonance structures;*
- *draw resonance structures of some covalent molecules and polyatomic ions;*
- *discuss the exceptions to the octet rule in covalent bonding;*
- *distinguish between polar and non–polar covalent molecules;*
- *describe the properties of covalent molecules;*
- *carry out an activity to investigate the effects of heat, electricity and some solvents on covalent compounds (naphthalene, graphite, iodine and ethanol);*
- *describe the valence shell electron pair repulsion theory (VSEPR);*
- *distinguish between the bonding pairs and non–bonding pairs of electrons;*
- *describe how electron pair arrangements and shapes of molecules can be predicted from the number of electron pairs;*
- *explain why double bonds and lone pairs cause deviations from ideal bond angles;*
- *explain the term dipole moment with the help of a diagram;*
- *describe the relationship between dipole moment and molecular geometry;*
- *describe how bond polarities and molecular shapes combine to give molecular polarity;*
- *predict the geometrical shapes of some simple molecules;*
- *construct models to represent shapes of some simple molecules;*
- *define intermolecular forces;*
- *name the different types of intermolecular forces;*
- *explain dipole - dipole interactions;*
- *give examples of dipole - dipole interaction;*
- *define hydrogen bonding;*
- *explain the effect of hydrogen bond on the properties of substances;*

- *give reason why H–bonding is stronger than ordinary dipole–dipole interaction;*
- *explain dispersion (London) forces;*
- *give examples of dispersion forces;*
- *predict the strength of intermolecular forces for a given pair of molecules.*

Forward Planning

Take ample preparation time to study covalent bonding and covalent compounds. This section discusses basic aspects of covalent bonding. The main ideas governing section 3.3 are writing Lewis structures, contributing structures (resonance structures), polar and nonpolar molecules, and properties of covalent compounds, molecular shapes, molecular polarity and intermolecular forces.

Teaching Aids

Ball and stick models are helpful if available. In case of shortages, you can prepare them from local materials. Alternatively you may encourage your students prepare models after giving them some methods of making such models. Diagrams sketched on flipcharts will help you also.

Subject Matter Presentation

Formation of Covalent Bonding

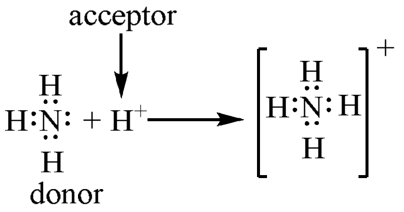
Discuss with the students that the heart of understanding chemistry is understanding what holds atoms together as molecules. One can also start to understand why, under certain conditions, old arrangements change in favor of new ones. Also, we can understand structure, and through structure, the nature and properties of the species under consideration.

Then bring the attention of the students to *Activity 3.8.* This activity is sufficient to facilitate discussions of the formation of covalent bonding. The suggested answers that follow may be helpful in harmonizing the discussion.

1. In an ionic bond, there is complete transfer of electrons, while in a covalent bond, there is sharing of electrons. In an ionic bond, the atoms are bound by attraction of opposite ions, whereas, in a covalent bond, atoms are bound by sharing electrons. In covalent bonding, the molecular geometry around each atom is determined by VSEPR rules, whereas, in ionic solids, the geometry follows maximum packing rules.
2. Pairs of valence electrons are shared. Once again, atoms try to complete an octet (or a duet in the case of hydrogen) in their valence shell. They share electrons in order to accomplish this.

3. A polar bond is formed when electrons are unequally shared between two atoms. Polar covalent bonding occurs because one atom has a stronger affinity for electrons than the other (yet not enough to pull the electrons away completely and form an ion). In a polar covalent bond, the bonding electrons spend a greater amount of time around the atom that has the stronger affinity for electrons. A good example of a polar covalent bond is the hydrogen-oxygen bond in the water molecule.

   A single covalent bond, in which both electrons in a shared pair come from the same atom, is known as a **coordinate covalent bond**. The donor atom provides both electrons to a co-ordinate covalent bond, and the acceptor atom accepts an electron pair for sharing a co-ordinate covalent bond. For example: a hydrogen ion unites with an ammonia molecule by a co-ordinate covalent bond to form the ammonium ion.



4. Aluminum tetrachloride ion, $AlCl_4^-$

   i. This compound is covalent.

   ii. Determine the total number of valence electrons available:

   One aluminum has 3 valence electrons

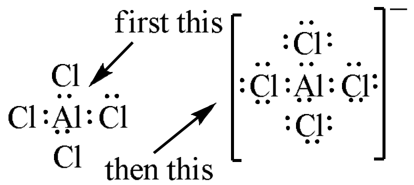
   Four chlorines, each with 7 valence electrons. Total = 28

   One negative charge means one electron. This means there are 3 + 28 + 1 = 32 valence electrons available, making 16 pairs.

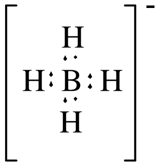
   iii. Organize the atoms so there is a central atom (usually the least electronegative) surrounded by ligand (outer) atoms. Hydrogen is never the central atom.

   Cl
   Cl Al Cl
   Cl

   iv. Determine a provisional electron distribution by arranging the electron pairs until all available pairs have been distributed:



Similarly, the Lewis dot structure for BH_4^- is



5. Both the length and strength of a bond change with bond order. See the table below.

Bond	Number of electrons	Bond order	Bond strength	Bond length
Single	2	1	Weakest	Longest
Double	4	2	Weaker	Average between the two
Triple	6	3	Strongest	Shortest

One analogy you can use for explaining bond strength is to think of atoms as a bundle of sticks. Then think of bonds as rubber bands. The rubber bands act as the force which holds the sticks together. As we increase the number of rubber bands, the sticks are squashed closer together and it takes more force to pull them apart.

In a molecule, as the number of electrons shared between two atoms increases, the bond order increases, the strength of the bond increases, and the distance between nuclei decreases.

Bond lengths are measured using X-ray and neutron diffraction, and bond strength by measuring the amount of energy needed to break a bond.

6. The two atoms approach and the following happens: the positive nucleus of each atom attracts the electron field of the approaching atom. The electron fields of both atoms repel, balancing the nuclear/electron attraction (the result is the final internuclear distance). The lone $3p^5$ electrons pair up, releasing some energy as the atoms become more stable.

Representation of Covalent Bond (Drawing Lewis Structures)

This part of chemistry is not totally new for Grade 11 students. Let them read the notes in the textbook and give them Exercises 3.4 and 3.5 as home work.

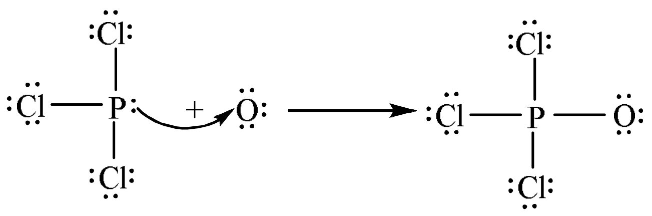
Coordinate Covalent Bonding

Activity 3.9 may be used to facilitate the discussion on this topic. The following suggested answers are given below.

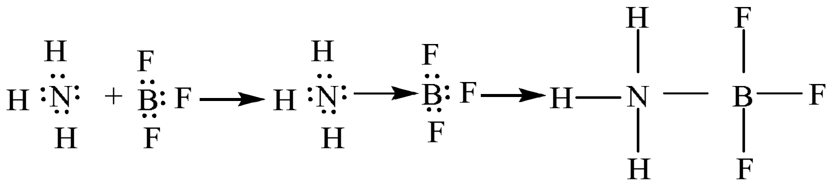
1. a. The nitrogen atom in the ammonia molecule has a lone pair of electrons, and it can provide both electrons to an electron-deficient atom or group of atoms.

   b. The oxygen atom can accept an electron pair from donor groups. This is the reason that the ammonia molecule and the oxygen atom participate readily in a coordinate covalent bond.

2. a.



   b. Sometimes, two or more stable molecules combine to form a molecular complex. In a molecular complex, the constituent molecules are held together by a 'coordinate bond'. One typical example involves the molecules of NH_3 and BF_3. The electron-dot structures of these molecules are:



ammonia-borontrifuloride
complex

The nitrogen atom has a complete octet around it, but the boron atom has only six electrons around it. The nitrogen atom therefore donates its lone pair of electrons to boron so that its atoms also acquire the octet. This one-sided sharing between N and B atoms gives rise to a coordinate bond. Coordinate bonds are involved in the formation of transition metal complexes known as coordination compounds.

3. All four hydrogen atoms and all four N-H bonds in the ammonium ion are found by experiment to be equivalent. Therefore, the one which is a coordinate covalent bond can't be identified.

Resonance Structures

Use the question in the students' textbook to start the lesson on resonance structure. Then direct *Activity 3.10*, which helps to provide further stabilization of the concept of resonance structure and resonance hybrid.

In chemistry, resonance is the appearance of delocalized electrons within certain molecules or polyatomic_ions, so that the bonding cannot be expressed by one single Lewis formula. A molecule or ion with such delocalized electrons is represented by several contributing structures (also called resonance structures)**.** Each contributing structure can be represented by a Lewis structure, with normal single, double or triple covalent bonds between every pair of adjacent atoms within the structure. These individual contributors cannot be observed in the actual resonance-stabilized molecule; resonance is not a rapidly-interconverting set of contributors. Several Lewis structures are used collectively to describe the actual molecular structure, which is an intermediate between the resonance forms. The intermediate form between different contributing structures is called a resonance hybrid. Contributing structures differ only in the position of electrons, not in the position of nuclei.

Exceptions to the octet rule

Remind the students that there are exceptions to the octet rule. Let them discuss in class the exceptions that fall into three main categories: less than octet, above octet, and odd-electron molecules. Then direct *Activity 3.11*. This activity characterizes the octet rules and violations of the octet rules. The answers suggested are as follows:

1. a, b, and d obey the octet rule.
2. The concept of the expanded octet occurs in any system that has an atom with more than four electron pairs attached to it. Most commonly, atoms will expand their octets to contain a total of five or six electron pairs, in total. In theory, it is possible to expand beyond these numbers. However the large amounts of negative charge concentrated in small volumes of space prevent those larger expanded octets from forming.

When an atom expands its octet, it does so by making use of empty d orbitals that are available in the valence level of the atom doing the expanding. Atoms that do not have empty valence level d orbitals cannot expand their octets. Second-period atoms cannot have more than eight valence electrons around the central atom, but atoms in the third (and below) period can. One example is SF_6, where all six of sulfur's valence electrons form bonds with fluorine. Other examples are AsF_5, XeF_6, and PF_5. The atom that expands its octet in a structure, will usually be located in the center of the structure and the system will not use any multiple bonds in attaching atoms to the central atom.

3. Phosphorus in PCl_5 and boron in BF_3 do not obey the octet rule

Polar and Non - Polar Covalent Molecules

Let the students read the notes in their textbook and try to recall what they have learned in Grade 9. Then facilitate discussions on this particular topic. Give them guiding questions to motivate their discussion and arguments during reading and discussions. Incorporate your suggestions in to *Activity 3.12* after the students have discussed on it thoroughly.

The answers suggested are as follows:

Solid substance	***Solvent***
Naphthalene	Carbon tetrachloride
Sodium chloride	Water
Iodine	Methyl benzene
Wax	Methyl benzene

Note: Naphthalene is insoluble in water because it is a fairly non-polar molecule (it is a hydrocarbon, and most hydrocarbons are non-polar) and water is a polar solvent. Wax is soluble in aromatic solvents, chloroform, esters, ethers and acetone. Iodine is more soluble in toluene than water. ***Like dissolves like!***

Experiment 3.5 is prepared to determine the melting point of naphthalene. Guide the students in setting up the apparatus in groups, as shown in the students' textbook (Figure 3.6). The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

Naphthalene ($C_{10}H_8$) is a simple molecule which is weakly bound together in crystals by van der Waals forces. So, the melting point of the crystal will be quite low. A pure solid melts at a temperature of 81^0C. Impurity depresses the melting point and usually causes gradual softening instead of sharp liquefaction. To melt sharply at its recognized (standard) melting point is good evidence for its purity for any known solid.

Experiment 3.5 is prepared to test the effect of heat, electricity some solvents on covalent compounds. The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis. The suggested answers are given below.

a. **Effect of heat:** if naphthalene and graphite are heated, their constituent particles, atoms, molecules, or ions acquire greater kinetic energy and vibrate more violently. Eventually, a point may come at which vibration overcomes the binding forces and the particles become mobile. The crystalline structure then collapses and a liquid state is reached in which the particles are free to move. The temperature, T°C, at which this occurs, is called the melting point of the solid and, at this temperature (and the prevailing pressure) the solid and liquid are in equilibrium.

The energy which must be supplied to convert the solid at T°C to liquid at the same temperature T^0C is known as *latent heat of fusion (melting)*. In the case of iodine, some molecules may break away from the crystal directly into the vapor phase. This is known as sublimation. When ethanol is heated, its molecules acquire increased kinetic energy. Therefore, the proportion of fast molecules increases, and evaporation is very rapid.

b. **Effect of electricity:** graphite is the only non – metal which conducts electricity. Naphthalene, iodine and ethanol do not conduct electricity, either in their solid state nor in molten state. No chemical reaction takes place at the electrodes.

c. **Solubility:** naphthalene and iodine are soluble in non-polar solvents; ethanol in water.

Activity 3.13 may be used for further discussion on the properties of covalent compounds. The following hints are given.

It is obvious that water is known as a universal solvent and it dissolves most substances. But non-polar molecules like methane do not dissolve in it for the mere reason that the intermolecular forces among them are not equivalent. Water molecules are strongly associated with each other through hydrogen bonding, whereas methane through weak London (dispersion) forces. There is hydrogen bonding in ammonia like water, and it is soluble in water. Again 'like dissolves like'.

Molecular Geometry

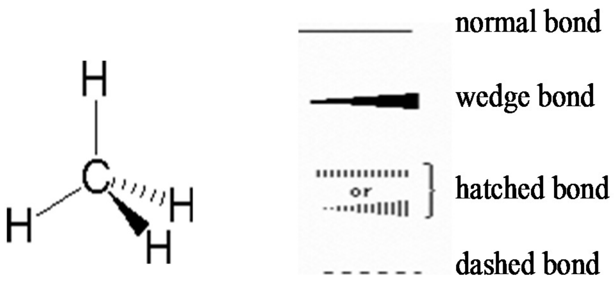
Inform the students that now, after spending a lot of time writing Lewis structures (in which symbols of elements are connected by line segments and surrounded by dots, stretching flat on the surface of the page), it is time to see the reality. This reality is a molecule extending in three dimensions throughout its tiny volume of space. Understanding this core idea of the topic to be discussed will make the teaching learning process simple and understandable. Notice that electron sets are repeatedly used to express pairs of electrons. Pairs are somehow confusing when it comes to multiple bonds and unpaired (single) electrons. The students must perform the exercises in the students' textbook and also some additional exercises clear understanding of the geometry of molecules and molecular ions.

Students should clearly know that the focus of this topic is the shape of molecules and its effect on the properties of a molecule.

Start this period's lesson by describing the valence shell electron pair repulsion (VSPER) model, which allows the conversion of a two-dimensional Lewis structure into three-dimensional geometry.

Inform the students that the three-dimensional shape or configuration of a molecule is an important characteristic. This shape is dependent on the preferred spatial orientation of covalent bonds to atoms having two or more bonding partners. Three-dimensional configurations are best viewed with the aid of models.

In order to represent such configurations on a two-dimensional surface (paper, blackboard or screen), use perspective drawings in which the direction of a bond is specified by the line connecting the bonded atoms. You can use different techniques of drawing so that students can view it from different angles. For instance, the following different lines can be used to illustrate the direction of bonds.



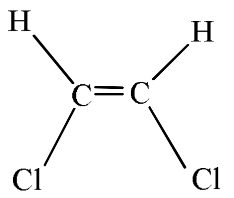
As defined in the above diagram, a simple straight line represents a bond lying approximately in the surface plane. The two bonds to H in the structure on the left are of this kind.

A wedge-shaped bond indicates a bond that reaches out toward the viewer. It is directed in front of the plane. It is thick end is toward the viewer. In contrast, a hatched bond is directed in back of the plane (away from the viewer).

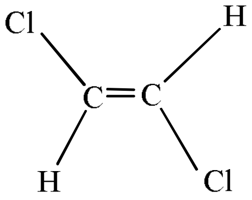
Discuss the basic principles and guidelines for applying this theory. Then engage students in *Activity 3.14* which may help to understand the nature and properties of dipoles and dipole moments.

The large electronegativity difference between hydrogen and oxygen obviously makes each H – O bond quite polar. But if the structure of water were linear, like many triatomic molecules, its dipole moment would have been zero and the molecule would be non-polar. Therefore, water is not linear like CO_2. This is the reason that its dipole moment is different from zero.

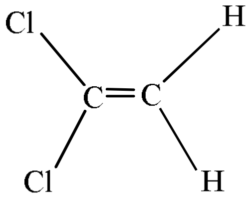
Activity 3.15 is used to enhance students'' understanding on bond polarity. Here is the suggested answer.



Z



X



Y

Compound Y has a non-zero dipole moment

Intermolecular Forces in Covalent Compounds

Discuss in detail the fact that the physical properties of substances are dependent on the intermolecular forces of attraction. Do not hesitate to explain that there are many types of intermolecular forces, but only three of them are mentioned in the students' textbook. The balance between these forces and the kinetic energy of the particles give rise to the properties of each state and to phase changes, which are the changes from one state (or form) to another. Facilitate and guide student participation on *Activity 3.16.* Use it to introduce the study of the vast intricate properties of molecules.

1. When we compare and contrast physical properties like boiling point and solubility, we usually consider comparable masses of substances under consideration.

   The intermolecular force in polar liquids is the dipole – dipole force (stronger), and in non-polar liquids it is London forces (weakest). It is easier to boil non-polar liquids than the polar ones.

2. The more polar a molecule – that is, the greater the dipole moment – the more pronounced is the effect of the dipole – dipole forces on physical properties. Hydrogen bonds (strongest) hold water molecules in a rigid structure keeping them very close to each other, whereas the weak London (dispersion) forces in methane cannot bring distant particles close to each other. This is one of the reasons that the particles of gases are far apart.

3. Benzoic acid and methanol are different from the other two-diethyl ether and decane. There is hydrogen bond in both benzoic acid and methanol, but there are two possible ends for hydrogen bonding in benzoic acid. The larger mass and the two possible ends in this acid make it solid at STP. Benzoic acid is a white crystalline solid. The rest are liquids under these conditions.

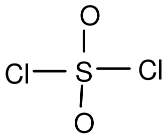
4. PCl_3

Activity 3.17 is used to further strengthen the discussion on the topic.

1. Dipole – dipole, hydrogen bond, London (dispersion) forces, etc.
2. The trans-isomer is non-polar, and the intermolecular force governing these types of molecules is the weak dispersion forces, whereas the cis-isomer is polar. The intermolecular force existing in polar covalent molecules is the strong dipole forces, which keeps the particles of the substance very close to each other. Therefore it exhibits high boiling points.
3. Compounds in which hydrogen is bonded to highly electronegative and smaller sized atoms such as nitrogen, oxygen, and fluorine form hydrogen bonding. These atoms exert a strong attraction for the bonding electron pair. The effect is very much like that of a bare proton at the end of the bond. It is represented as X—H---Y, where X is the atom to which the hydrogen is actually normally bonded, and Y is the atom to which the hydrogen is bonded. It has been found experimentally that hydrogen bonds of significant strength are formed when the atoms X and Y are either N, F, or O. If either X or Y is another element (e.g., Cl), much weaker hydrogen bonds are formed.
4. Polarizability of large molecules, molecular mass and molecular geometry are the essential factors to consider.
5. The hydrogen bond in ice holds the water molecules in a rigid but open structure and more volume is occupied. As ice melts, some of the hydrogen bonds are broken. Water molecules move into some of the holes in the rigid structure of ice. This means that liquid water at 0^0C is more dense than ice.

Additional Notes

In order to use Lewis formulas effectively, the students should be familiar with the set of rules stated in the text book. Alternatively, you may make the students practice the following simple step-by-step method for drawing Lewis structures. Take for example, SO_2Cl_2 to apply the rules.



$$\mathrm{Cl-\overset{\overset{\displaystyle O}{|}}{\underset{\underset{\displaystyle O}{|}}{S}}-Cl}$$

1. Decide which atoms are bonded:
2. Calculate the total number of valence electrons:

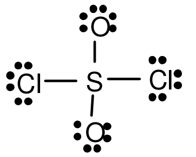
S: 1× 6 = 6

O: 2 × 6 = 12

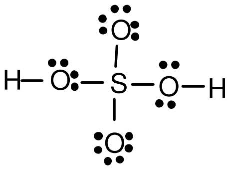
Cl: 2 × 7 = <u>14</u>

32

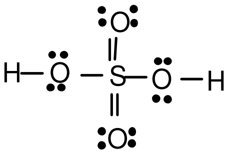
3. Place six electrons on each terminal atom to make it octet.



4. All the 32 electrons are now placed. There are no remaining electrons. If the central atom had been below octet, you would form double bonds, or triple bonds. If electrons were remaining, you could put them on the central atom. Students must master the process by practicing it rather than by memorizing the rules. You should help your students to note that Lewis structures are only theoretical explanations and predictions based on observations. Sometimes what is expected from the rules and what is actually observed are different. For example, sulfuric acid is expected to exhibit the structure depicted below, based on the Lewis rules:



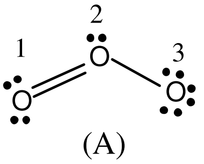
But the observed (actual) structure of H_2SO_4 is

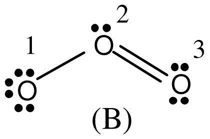


This can be explained in terms of exceptions to the octet rule-those with the expanded octet. Sulfur has a valence shell that has 3s, 3p, and 3d subshells, which can accommodate more than eight electrons.

A formal charge is a hypothetical charge the atom has in the molecule or ion.

Formal charge of atom = n$\underline{o}$. of valence electrons - (n$\underline{o}$. of unshared electrons e^- + $\frac{1}{2}$ n$\underline{o}$ of shared valence e^- because an atom is assigned all of its unshared valence electrons and half of its shared valence electrons. Of course, the formal charges must sum up to the actual charge on the species: zero for a molecule and the ionic charge for an ion. In the two O_3 resonance forms in the students' textbook, the formal charge can be calculated as follows:


(A)


(B)

The numbers (1), (2) and (3) are arbitrarily assigned to indicate the three oxygen atoms in the molecule.

The formal charges in the two forms are:

Structure (A)

$$\mathbf{O}^1\ [6 - 4 - \frac{1}{2}(4)] = 0$$

$$\mathbf{O}^2\ [6 - 2 - \frac{1}{2}(6)] = +1$$

$$\mathbf{O}^3\ [6 - 6 - \frac{1}{2}(2)] = -1$$

Structure (B)

$$\mathbf{O}^1\ [6 - 6 - \frac{1}{2}(2)] = -1$$

$$\mathbf{O}^2\ [6 - 2 - \frac{1}{2}(6)] = +1$$

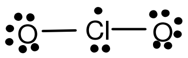
$$\mathbf{O}^3\ [6 - 4 - \frac{1}{2}(4)] = 0$$

Structures A and B are symmetrical – each has the same formal charges but on different atoms of oxygen – so they contribute equally to the resonance hybrid.

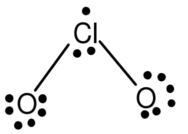
Three general points govern the most reasonable (plausible) Lewis structure based on formal charges:

1. Smaller formal charges (whether positive or negative or 0) are preferable to large ones.
2. Like charges on adjacent atoms are not desirable.
3. A more negative formal charge should reside on a more electronegative atom.

Explain to your students that the VSEPR theory is based on the idea that valence shell electrons pairs need to stay as far apart from each other as possible. VSEPR assume that the geometry of a molecule depends only upon electron-electron interactions. Before applying the VSEPR method of predicting molecular geometry, it is necessary to write down a Lewis electron-dot structure for the molecule under consideration. A slight complication is seen in the geometry of a molecule when there are unpaired electrons. This is well-illustrated for ClO_2, in which there is a non-integral number of electron pairs. Since any orbital can accommodate 0, 1, or 2 electrons, the lone-pair orbital is only half-filled, which demands less space. The suggested Lewis structure for ClO_2 is



It is considered as if there are two lone-pairs and two bond pairs around the central atom, chlorine, making the electron set arrangement tetrahedral and the geometry of the molecule angular or bent.



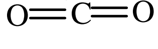
The valence shell electron pair repulsion theory (VSEPR) is very useful in predicting the structures of molecules. The predictions of the model are basically contained in the students' textbook. The VSEPR model is applied to predicting the structures of molecules or ions containing multiple bonds. What would you have predicted for the shapes of XeF_2 and XeF_4? You can simply predict that:

XeF_2 is an AX_2E_3 type, with two bonding and three lone pairs around the central atom and a linear geometry of a molecule and a trigonal bipyramidal electron set arrangement.

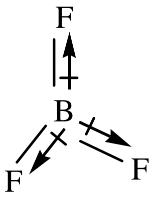
XeF_4 is an AX_4E_2 type, with four bonding and two lone pairs of electrons around the central atom. Therefore, the shape of the molecule is square planar and the electron arrangement is octahedral.

The dipole moment is defined as the product of the charge at either end of the dipole, Q, times the distance, r, between the charges: $u = Qr$. Thus, the dipole moment increases as the quantity of charge which is separated increases, or as the distance between the positive and negative centers increases.

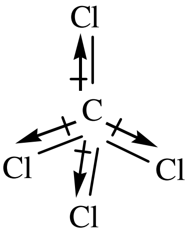
The dipole moment of a molecule containing more than two atoms depends both on bond polarity and molecular geometry. To have a dipole moment, polyatomic molecules must have polar bonds. However, even if polar bonds are present, the molecule itself might not have a dipole moment if the bonds are arranged that so their polarities cancel. This situation is found in the linear CO_2 molecule



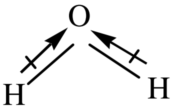
The diagram shows a polar bond but a nonpolar molecule. Other examples are H_2O, NH_3, BF_3, CCl_4, and CH_3Cl.



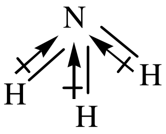
Non - polar molecule



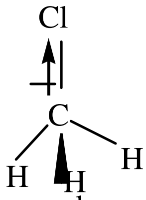
Non - polar



polar



polar



polar

Assessment

You are required to assess students' active involvement in each of the activities. At the end of each lesson, ask short answer questions as give implement your lesson plan. Your student-performance record list is helpful in all your assessments. Don't forget to show application for students working above the minimum requirement level. You are required to assist those working below the minimum requirement level.

Additional Questions

* 1. Which of these molecules has an atom with an expanded octet?

a. HCl b. $AsCl_5$ c. ICl d. NCl_3 e. Cl_2

2. Among the following molecules with highest dipole moment is:

a. CH_3Cl b. CH_2Cl_2 c. $CHCl_3$ d. CCl_4

* 3. Hydrogen bonding is not present in:

a. Glycerol b. Water

c. Hydrogen sulfide d. Hydrogen fluoride

4. Which of the following hydrogen halide is most volatile?

a. HF b. HCl c. HBr d. HI

5. The dipole moment is highest for:

a. Trans-2-butene b. 1,3- dimethyl benzene

c. Acetophenone d. Ethanol

* 6. Which of the following contains both covalent and ionic:

a. NH_4Cl b. CCl_4 c. $MgCl_2$ d. KCl

7. Which of the following has a dipole moment?

a. CCl_4 b. NH_3 c. CS_2 d. BF_3

* 8. What is the octet rule in chemistry?

*9. What does delocalized mean?

10. How do hydrogen bonds affect heat capacity?

*11. Compare a polar covalent molecule and a non-polar covalent molecule in terms of electrical conductivity.

12. What is a dipole moment?

13. When are non-polar covalent bonds formed?

14. Is a covalent bond stronger than a hydrogen bond?

*15. What sort of covalent molecule is water?

*16. What is a non-polar compound?

*17. List three common properties of covalent compounds.

18. Why do covalent bonds have an odor?

19. When is a compound called a molecule?

Answers to Additional Questions

1. b
2. a
3. c
4. b
5. d
6. a
7. b
8. The octet rule is a simple chemical rule of thumb that states that atoms tend to combine in such a way that they each have eight electrons in their valence shell.
9. The term 'delocalized' refers to an electron which is not 'attached' to a particular atom.
10. Hydrogen bonds such as those in water are strong. Therefore, it takes more energy (heat) to bring it to a boiling point (break or weaken the force).
11. Neither of them are conductors of electricity (whether in solid form, or dissolved in water).
12. Dipole moment is the measure of polarity of a polar covalent bond.
13. When the two atoms forming the bond have similar or equal electronegativity (the ability of an atom to pull electrons towards itself is equal).
14. Yes, a covalent bond is stronger than a hydrogen bond.
15. It is a polar covalent molecule.

16. There is no non polar compound. However, you can talk about a non-polar molecule. It is a molecule with equal distribution of electrons around each atom.
17. Low melting and boiling points, do not conduct electricity in water, soft (compared to ionic compounds).
18. Odor is not typical of covalent bonds. It is a characteristic of covalent compounds.
19. A molecule is formed when two or more atoms join together chemically by covalent bonding. A compound is a molecule that contains at least two different elements. Covalent compounds are molecules, but not all molecules are covalent compounds.

    Molecular hydrogen (H_2), molecular oxygen (O_2) and molecular nitrogen (N_2) are not compounds because each is composed of a single element. Water (H_2O), carbon dioxide (CO_2) and methane (CH_4) are compounds because each is made from more than one element. The smallest bit of each of these substances would be referred to as a molecule.

    For example, a single molecule of molecular hydrogen is made from two atoms of hydrogen, while a single molecule of water is made from two atoms of hydrogen and one atom of oxygen.

Answers to Exercise 3.4

a. CO_2 has $(1 \times 4) + (2 \times 6) = 16$

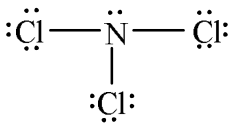
b. SO_4^{2-} has $(1 \times 6) + (4 \times 6) + 2 = 32$

c. NH_4^+ has $(1 \times 5) + (4 \times 1) - 1 = 8$

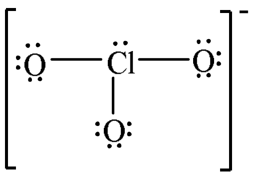
d. N_2O_4 has $(2 \times 5) + (4 \times 6) = 34$

Answers to Exercise 3.5

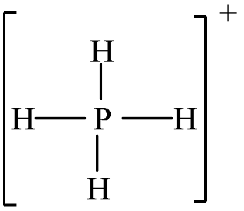
1. a



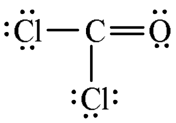
b.



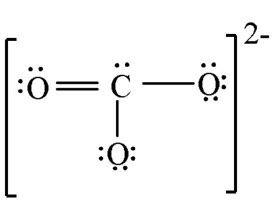
b.

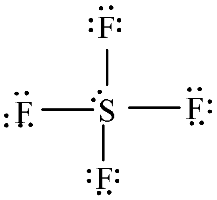


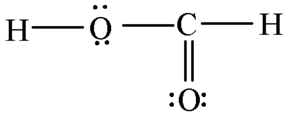
d.



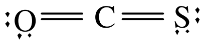
2.





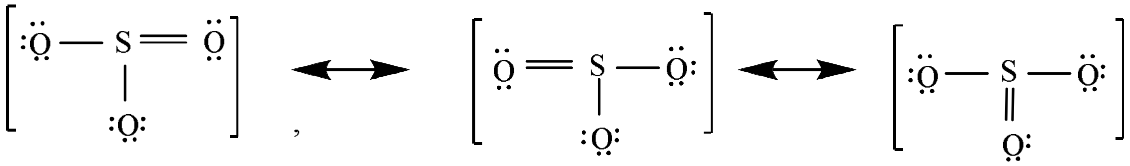


3. F and H

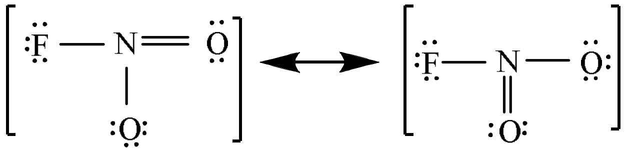
 4.

Answers to Exercise 3.6

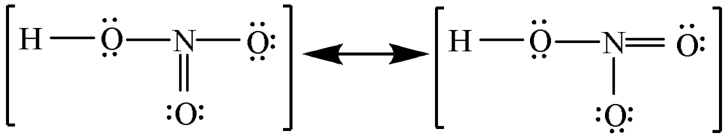
1.



2. a.

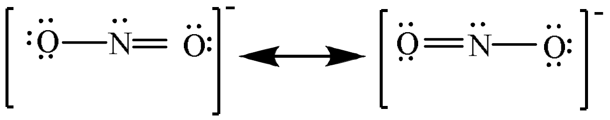


b.



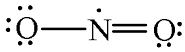
c. No resonance from.

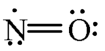
d.



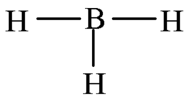
Answers to Exercise 3.7

1.



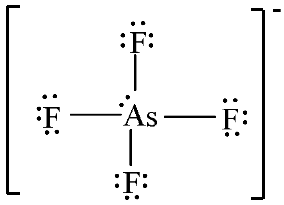


2. a.



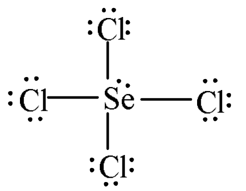
(Below octet)

b.



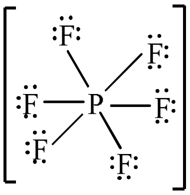
(Above octet)

c.



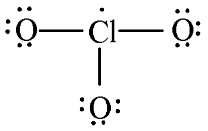
(Above octet)

d.



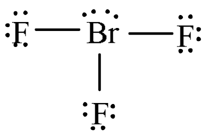
(Above octet)

e.



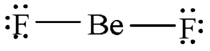
(Odd number of electrons)

f.



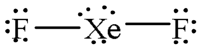
(Above octet)

g.



(Below octet)

h.



(Above octet)

Answers to Exercise 3.8

From their electronegativity data, one can predict that there is a bond moment in both CO_2 and BCl_3. But from the molecular geometry of these molecules, it can be predicted that the bond moments cancel to give a non-polar molecule to produce a zero resultant dipole moment.

Answers to Exercise 3.9

a. Tetrahedral, 109° b. Linear, 180° c. Angular (bent), 104°

Answers to Exercise 3.10

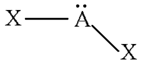
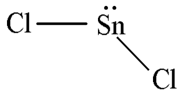
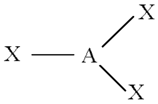
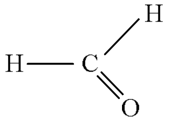
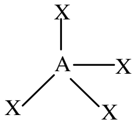
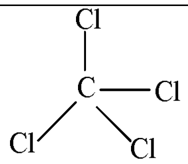
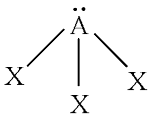
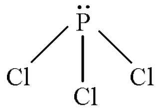
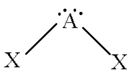
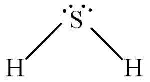
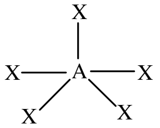
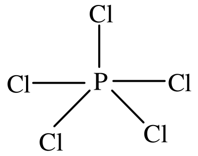
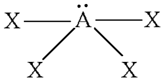
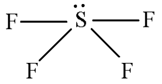
1. When there is no lone pair of electron(s) on the central atom.

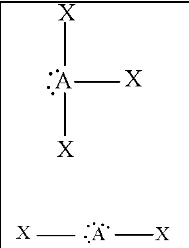
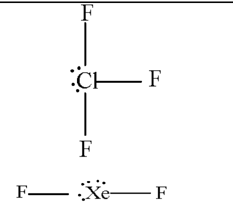
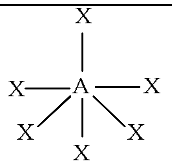
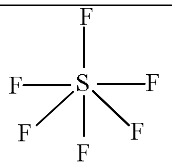
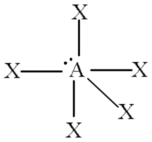
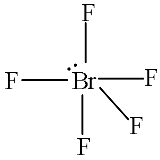
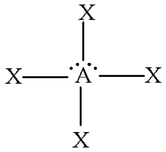
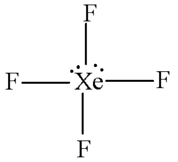
2. $OF_2 < NF_3 < CF_4 < BF_3 < BeF_2$

3. $PCl_5(g)$ (AX_5 type) has a trigonal bipyramidal molecular shape. In the solid phase, PCl_4^+ (AX_4 type) is tetrahedral, and PCl_6^- (AX_6 type) is seesaw in shape.

4. By the vector sum of bond moments

5.

No of electron pairs	***Molecular formula***	***Molecular geometry***	***Electron set geometry***	***Sketch***	***Example***
2	AX_2	Linear	Linear	X–A–X	Cl–Be–Cl
3	AX_2E	Angular	Trigonal planar		
	AX_3	Trigonal Planar	Trigonal Planar		
4	AX_4	Tetrahedral	Tetrahedral		
	AX_3E	Trigonal pyramidal	Tetrahedral		
	AX_2E_2	Angular	Tetrahedral		
5	AX_5	Trigonal bipyramidal	Trigonal bipyramidal		
	AX_4E	Seesaw	Trigonal bipyramidal		
	AX_3E_2	T-shape	Trigonal bipyramidal		

	AX_2E_3	Linear	Trigonal bipyramidal		
6	AX_6	Octahedral	octahedral		
	AX_5E	Square pyramidal	octahedral		
	AX_4E_2	Square planar	octahedral		

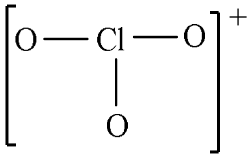
6. a. Trigonal bipyramid b. square pyramidal

7. a. Trigonal planar (120^0) b. Trigonal planar (120^0)

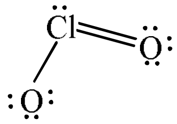
c. Angular (bent) (104^0) d. Linear (180^0)

e. Trigonal bipyramidal (90^0, 120^0) f. T – shape (90^0)

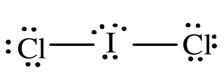
8. a.



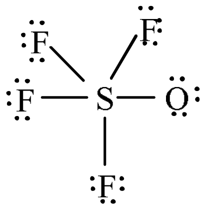
b.



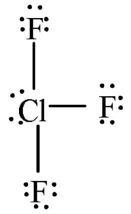
c.



d.



e.



Answers to Exercise 3.11

1. In b and c.

2. a. ethanol (hydrogen bonding) b. hexane (Dispersion forces)

c. magnesium chloride (ionic) d. methyl amine (hydrogen bonding)

3. a. dipole – dipole b. Hydrogen bonding

c. London e. Dipole – dipole

4. Intramolecular forces are chemical bonds which are stronger than any of the intermolecular forces.

3.4 METALLIC BONDING

Periods allotted: 2 periods

Competencies

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

- *explain how metallic bond is formed;*
- *explain the properties of metals related to the concept of bonding;*
- *carryout an activity to investigate the conductivity, malleabily and ductility of some metals and nonmetals (Al, Cu, Fe, Sn, Zn, S, C–charcoal, C–graphite and Si).*

Forward Planning

Take enough preparation time to study metallic bonding. Plan to use some illustrative methods of teaching so that the students will easily grasp the idea of bonding type and the properties of metals. You may supply the students with different substances for than group assignments. Thus, make the necessary materials ready ahead of time.

Teaching Aids

Diagrams sketched on flipcharts will help. Also, metals of different types and ionic crystals can be used for comparison. The materials required for the group assignment given in this section are also teaching aids.

Subject Matter Presentation

Formation of Metallic Bonding

Start this topic with the in-text question:

Is a metal made up of atoms or ions?

Special attention should be given here to the discussion of metallic bonding. It is usually confusing and many students are under the impression that it is a bond between two or more different metals. Try to make the concept clear through discussions on matters related to metallic bonding. Note that any successful theory of metallic bonding should explain the bonding of a large number of identical (not similar) atoms in a pure metal and the bonding between widely different metal atoms in alloys. You will not have Sodium (Na) and Potassium (K) interacting with each other! Explain that the mobility of electrons moving through the metal is similar to the free movements of molecules of a gas.

Activity 3.18 may be used to facilitate discussion on the formation of metallic bonding. The following suggested answers are given below.

1. Metals are very reactive and form compounds with other elements quite easily. Sodium (Na) and potassium (K) are some of the most reactive metals.

   Some of the physical properties which characterize metals are:

   a. They have high electrical and thermal conductivity.
   b. They have bright lustrous appearance.
   c. They are malleable and ductile.
   d. They are very cool and shiny.
2. Metallic bond.
3. Electrons which are not confined to a particular atom (mobile electrons).
4. The model which represents the electronic structure of metals.
5. The alkaline earth metals have the valence shell electron configuration ns^2. Therefore, their atoms show a tendency to give up two electrons to form 2+ ions. which makes them isoelectronic with the nearest noble gas.

Properties of Metals Related to the Concept of Bonding

As you facilitate this lesson have encourage your students read and discuss in groups. Use your presentation to the students to refine their understanding of the subjects. Use the suggested activity to stabilize their understanding.

Let the students discuss the physical properties of metals such as conductivity, malleability, ductility, luster, hardness, etc. For example, to show that metals are hard and do not break up on hammering, use a metal like copper or aluminum foil and an ionic compound NaCl and a hammer. Ensure that the students understand that metals are malleable and ductile, and good conductors of heat and electricity, because of the nature of their bonding, particularly as compared to ionic solids.

Activity 3.19 may be used to facilitate discussion on the properties of metals. The following suggested answers are given below.

Your left hand feels colder because of the high heat conductivity of metals. Heat flows from your hand to the metal object.

Activity 3.20 is used to facilitate further discussion on the properties of metals. The following suggested answers are given below.

1. $O_2 < Br_2 < K < Mg$ because oxygen and bromine are non-metals with covalent bonds. The dominant intermolecular force governing these molecules is the London (dispersion) force. The strength of London forces increases with molecular mass. Br_2 – liquid at room temperature is larger in mass than O_2 – gas at room temperature. Correspondingly, K and Mg are metals with 1+ and 2+ oxidation states which will contribute 1 and 2 electrons, respectively, for metallic bonding. The higher the number of electrons contributed, the stronger is the metallic bond and the higher are the melting and boiling points.
2. a. They have high electrical and thermal conductivity.
    a. They have bright lustrous appearance.
    b. They are malleable and ductile.
    c. They are hard.
3. a. The boiling point is higher due to the large amount of energy necessary to separate the metal ions from each other in the electron "sea".
    b. Be has two valence electrons per atom, compared with Li which has one. The metallic bond is therefore stronger for Be.
    c. K is a larger atom than Na. Therefore its electrons are held more loosely, and thus its metallic bond strength is weaker.
4. a. Li is more metallic than Be. Metallic character decreases from left to right across a period.
    b. Na is more metallic. Metallic character increases down a group.
    c. Pb is more metallic for the same reason given in b.
    d. Al is more metallic.

Group Assignment

Organize students into groups and supply them with the materials required for the assignment. Guide them as they perform the suggested activity to the extent of arranging set ups and different connections.

Assessment

Assess students by asking them to compare and contrast the three basic types of chemical bonding.

You may ask the students to explain conductors, insulators and semiconductors, using the bonding theories. You may ask them also to state the physical properties of metals and to describe how the bonding theories account for them. Account for the electrical conductivity of metals and explain why they are not brittle.

Additional Exercises

* 1. Why was bronze used before iron?

2. How do you distinguish a metallic bond from an ionic bond or a covalent bond?

Answers to Additional Questions

1. Bronze was used before iron because it was easier to mine, find and mold to required uses.
2. Metallic bonding is the bonding within the atoms of a metal. It is the hybrid form of ionic and covalent bonding.

3.5 CHEMICAL BONDING THEORIES

Period Allotted: 8 periods

Competencies

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

- *name the two chemical bond theories;*
- *explain the valence bond theory;*
- *distinguish the Lewis model and the valence bond model;*
- *discuss the overlapping of orbitals in covalent bond formation;*
- *explain hybridization;*
- *show the process of hybridization involved in some covalent molecules;*
- *draw hybridization diagram for the formation of sp,* sp_2, sp_3, sp_3d *and* sp_3d_2 *hybrids;*
- *suggest the kind of hybrid orbitals on the basis of the electron structure of the central atom;*

- *predict the geometrical shapes of some simple molecules on the bases of hybridization and the nature of electron pairs;*
- *discuss the hybridization involved in compounds containing multiple bonds;*
- *explain bond length and bond strength;*
- *explain molecular orbital theory;*
- *describe molecular orbital using atomic orbitals;*
- *describe bonding and anti–bonding orbitals;*
- *draw molecular orbital energy level diagrams for mononuclear diatomic molecules;*
- *write the electron configuration of simple molecules using molecular orbital model;*
- *define bond order and determine the bond order of some simple molecules and ions;*
- *determine the stability of a molecule or an ion using its bond order; and*
- *predict magnetic properties of molecule.*

Forward Planning

Be sure to explain that MOT and VBT are complementary, and therefore one of the other should be used when it has advantages over the other. Students could get frustrated, as they would like one to work best all the time. The thing to keep in mind throughout discussions of the different approaches is that each theory is attempting to describe the same thing, and that each in its own way succeeds to a certain degree. Take note that the topics VBT and MOT are vast and may be abstract. Undoubtedly you will be working against time to get the lesson completed within the allotted periods.

Teaching Aids

Models and diagrams sketched on flipcharts may help.

Subject Matter Presentation

Valence Bond Theory (VBT)

Students should have a clear image of the VBT and be able to understand, compare and contrast at least the three theories: Lewis theory, VSEPR theory and VBT. Ask the students if all scientific models have limitations. As we all know, scientific models are simplifications of reality.

Note that, no explicit attempt was made in the students' textbook to convey the idea of the shape of a molecule through Lewis structure until we discussed VSEPR theory. This approach emphasizes that the prediction of molecular shapes comes not from Lewis theory but from VSEPR theory.

Discuss the limitations of Lewis and VSEPR. Explain the basic principles of the valence bond model.

Explain, using examples, that many of the aspects that could not be explained by the simple Lewis and VSEPR models are well elucidated by the VB.

Try to find an analogy to describe the hybridization of atomic orbitals for example, the blending of coffee and milk. Be creative as you plan your explanations of the concept. The exercises and examples given in the students' text are helpful stabilizing the lesson.

Molecular Orbital Theory (MOT)

Tell the students about MOT, and explain that VBT has limitations as well, because it does not adequately explain the magnetic or spectral properties of molecules. The stability of certain molecules can best be described using MOT theory. The importance of electron delocalization is considered less important in VBT. Molecular orbital theory (MOT) is important for dealing with these phenomena,

Tell the students that the order of filling molecular orbitals is dictated by the rules of all orbital filling – the Aufbau, Pauli exclusion and Hund's principles.

Combination of Atomic Orbitals

In presenting this lesson, emphasize the following main points during the teaching-learning process. In the molecular orbital theory, the valence electrons are thought to be associated with all the nuclei considered. To do this, atomic orbitals from different atoms must be combined, (added together) to produce molecular orbitals.

Explain that electrons may be considered either as particles or as waves. An electron in an atom may therefore be described as occupying an atomic orbital.

Electron Configuration of Diatomic Molecules

It is important that the students have a complete grasp of the energy order of molecular orbitals in homonuclear diatomic molecules. Let the students notice that the molecular orbitals are filled with electrons from lowest to highest energy. Electrons are not paired until all orbitals of the same energy are half-filled. Also, let the students note that molecular orbital theory defines bond orders.

Additional Notes

The molecular orbital theory looks at the molecule as a collection of positive nuclei surrounded by electrons occupying sets of molecular orbitals. Two atomic orbitals combine to form bonding and antibonding molecular orbitals. Generally, the bonding orbitals are identified as sigma (σ) or pi (π), while the corresponding anti-bonding orbitals marked with an asterisk (*) are as sigma antibonding (σ^*) or pi antibonding (π^*).

Electrons in a bonding orbital spend most of their time between two nuclei and act rather like a sticky material holding the atoms together. Electrons in an antibonding orbital spend most of their time outside the region between the two nuclei. The nuclear charges are not screened from each other by the negative charge of electrons, and hence they repel each other.

Assessment

You may use the suggested exercises as class work or homework, as required. You are also advised to prepare additional questions for students so you can assess their mastery of the concepts presented.

Assess each student's work to determine whether they have achieved the minimum required level or not and record the results on the record list. You can use the exercises and activities in the students' text or develop additional exercises to evaluate their level of understanding.

Additional Questions

1. Explain the basic principles of the valence bond model.
2. The molecules/species having an unpaired electron are:

   a. F_2 b. CN^- c. NO d. O_2
3. Which of the following species is paramagnetic?

   a. O_2^- b. CN^- c. CO d. NO^+
4. Which electron is most electronegative?

   a. Sp^3 hybridized. b. Sp hybridized.

   c. Sp^2 hybridized. d. None of these.
5. How many sigma and pi (π) bonds are present in toluene?

   a. 3π and 8 sigma bonds b. 3 π and 16 sigma bonds

   c. 3 π and 10 sigma bonds d. 6 π and 3 sigma bonds

6. Which of following is not paramagnetic?

   a. N_2^+ b. CO c. O_2^- d. NO

7. The number of pi(π) - bonds in naphthalene is :

   a. 6 b. 3 c. 4 d. 5

Answers to Additional Questions

1. The basic principle of the valance bond model is that a covalent bond forms when the orbital from two atoms overlap and a pair of electrons occupies the region between the nuclei. Therefore, it bases molecular shape on orbital characteristics.

2.	c and d	3.	a
4.	b	5.	c
6.	b	7.	d

Answers to Exercise 3.12

1. Carbon uses the s and p subshells from the second principal energy level, whereas silicon uses those from the third principal quantum number. The hybridized orbital in both cases is sp^3.

2. a. Two sigma bonds comprise two single bonds.

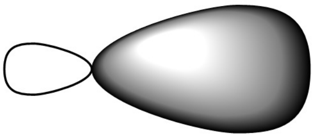
   b. A triple bond consists of one σ and two π bonds.

   d. a π bond consists of a pair of electrons.

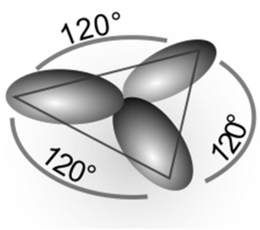
   e. end – to – end overlap results in a bond with electron density along with the internuclear axis.

Answers to Exercise 3.13

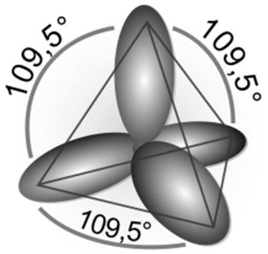
1. sp, sp^2, sp^3; linear, trigonal planar, and tetrahedral shapes, respectively.



sp hybridized



sp^2 hybridized



sp^3 hybridized

2. sp, linear geometry.

Answers to Exercise 3.14

1. a. sp^2, trigonal planar.

   b. Sulfur has 6 valence electrons, and so does oxygen. Therefore, SO_2, we're dealing with 18 valence electrons, total.

   First, draw single bonds between the S and each O, then fill out an octet for each O. This uses 16 electrons. The other two should be placed as a lone pair on the S.

   Finally, move a lone pair on one of the O's so that it forms a second bond with the S. This completes an octet for all atoms.

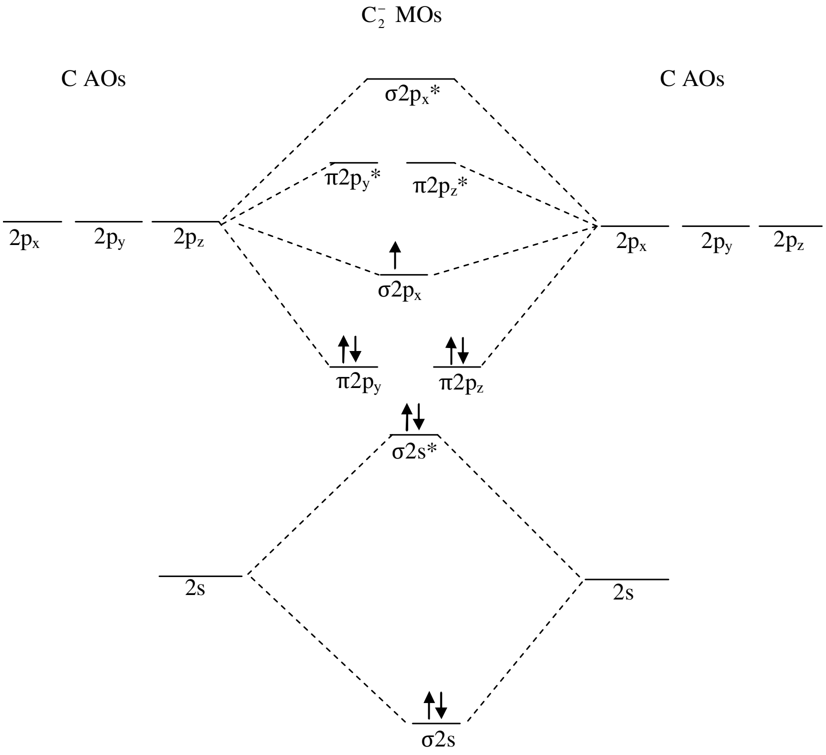
   Now count electron groups around the S: there's a double bond, a single bond, and a lone pair.... that makes 3. So the S in SO_2 is sp^2 hybridized with trigonal planar structure.

2. sp, linear.

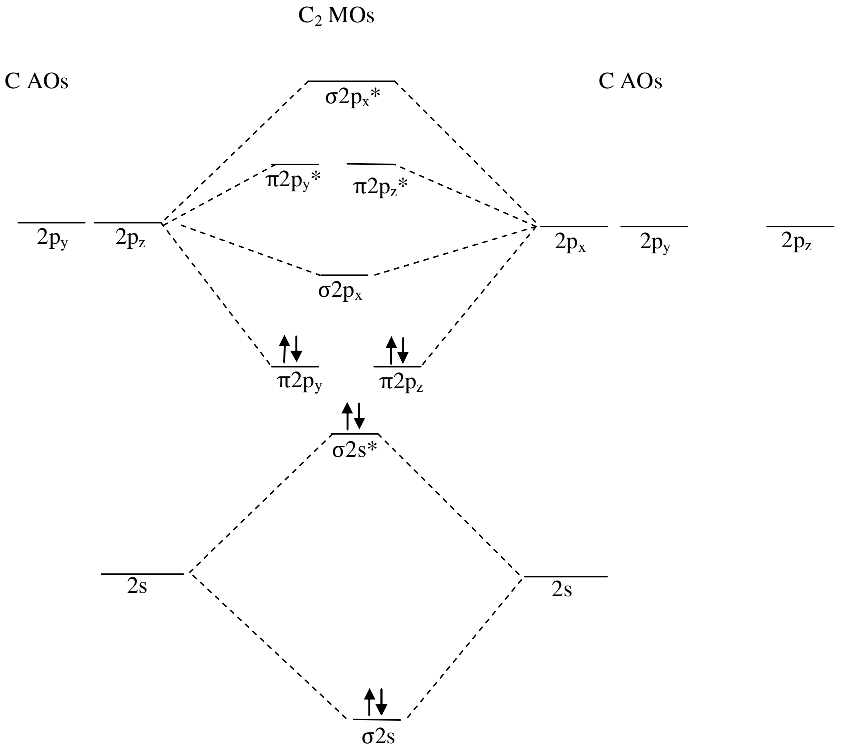
3. Nitrate ion exhibits resonance structure. The true structure is the hybrid of the resonance structures, and the bond length is about 129 pm, the bond angle = 120^0 trigonal planar in geometry, three sigma and one pi bond.

Answers to Exercise 3.15

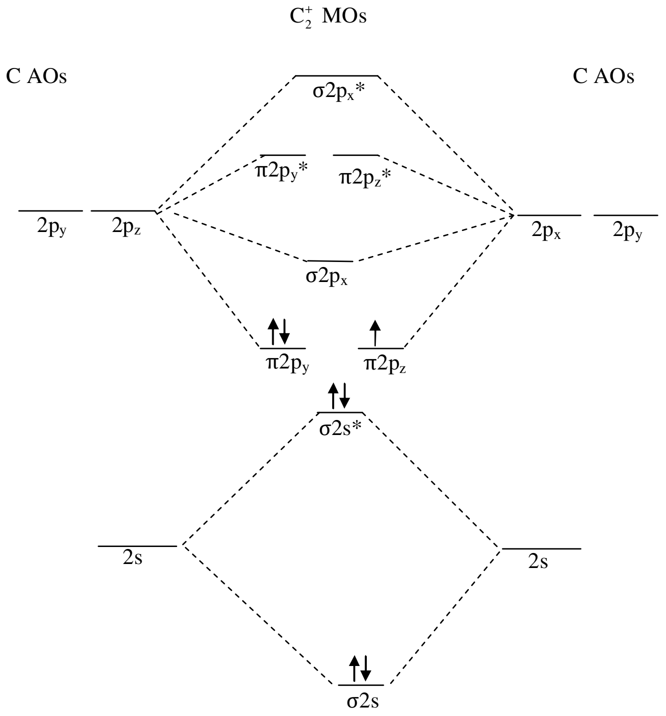
1. a.



b.


C_2 MOs
C AOs
C AOs
$\sigma 2p_x^*$
$\pi 2p_y^*$
$\pi 2p_z^*$
$2p_y$
$2p_z$
$2p_x$
$2p_y$
$2p_z$
$\sigma 2p_x$
$\pi 2p_y$
$\pi 2p_z$
$\sigma 2s^*$
2s
2s
$\sigma 2s$

c.


C_2^+ MOs
C AOs
C AOs
$\sigma 2p_x^*$
$\pi 2p_y^*$
$\pi 2p_z^*$
$2p_y$
$2p_z$
$2p_x$
$2p_y$
$\sigma 2p_x$
$\pi 2p_y$
$\pi 2p_z$
$\sigma 2s^*$
2s
2s
$\sigma 2s$

Answers to Exercise 3.16

a. $\sigma 1s^2\sigma^*1s^2$. Does not exist.

b. $\sigma 1s^2\sigma^*1s^2\ \sigma 2s^2\sigma^*2s^2$. Does not exist.

c. $\sigma 1s^2\sigma^*1s^2\sigma 2s^2\sigma^*2s^2 \begin{Bmatrix} \pi_2 p_y^1 \\ \pi_2 p_z^1 \end{Bmatrix}$. Exists.

d. $\sigma 1s^2\sigma^*1s^2\sigma 2s^2\sigma^*2s^2 \begin{Bmatrix} \pi_2 p_y^2 \\ \pi_2 p_z^2 \end{Bmatrix}$. Exists

e. $\sigma 1s^2\sigma^*1s^2\sigma 2s^2\sigma^*2s^2\sigma 2p_x^{\ 2} \begin{Bmatrix} \pi_2 p_y^2 \\ \pi_2 p_z^2 \end{Bmatrix}$. Exists

Answers to Exercise 3.17

1. Bond order = 3 (CN^-) = 2.5 (CN) = 2 (CN^+)
2. None is paramagnetic.

3.6 TYPES OF CRYSTALS

Periods allotted: 1 period

Competencies

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

- *define crystal;*
- *name the four types of crystalline solids and give examples;*
- *mention the types of attractive forces that exist within each types of crystalline solids;*
- *describe the properties of each type of crystalline solids;*
- *build a model of sodium chloride crystal structure.*

Forward Planning

Take ample time of preparation to discuss about crystals and the properties related to intermolecular forces.

Teaching Aids

Models of some simple structures

Subject Matter Presentation

As you implement this lesson, you may assign your students to read and discuss in groups. Students should know that types of crystals are distinguished from each other by the type of intermolecular forces between the individual molecular particles or ions, which in turn determine physical properties like melting and boiling points and solubility in different solvents. Students should be asked to construct a model of the NaCl structure from balls and sticks or other locally available materials. Use the suggested group assignment and exercises at the end to stabilize their understanding.

Assessment

Asking short-answer questions will help you assess students at the end of the lesson.

Answers to Exercises

Exercise 3.18

a. -------------- e because Fe is a heavy metal.
b. -------------- d I_2 is non-polar covalent.
c. -------------- c there is hydrogen bonding in H_2O.
d. -------------- b SiO_2, a covalent network crystal.
e. -------------- a metallic bonding.

Answers to Review Questions and problems

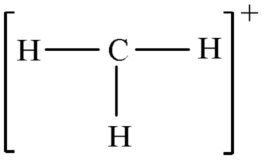
1. C
2. C
3. C
4. C
5. A
6. A
7. B
8. B
9. B
10. A
11. D
12. C
13. A
14. C
15. A
16. D
17. D
18. A
19. D
20. D

Short-answer questions

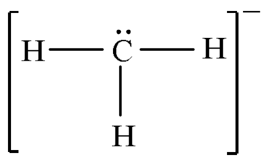
21. a. Potassium bromide
b. Aluminum iodide
c. Lithium nitride
d. Beryllium oxide
e. Barium sulfide

22. a. Tetrahedral
b. Trigonal pyramidal
c. Angular (bent)
d. Trigonal pyramidal
e. Tetrahedral
f. Trigonal bipyramidal

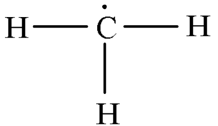
23.



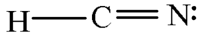
trigonal planar



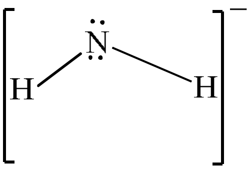
trigonal pyramid

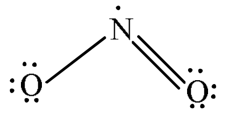


trigonal pyramid

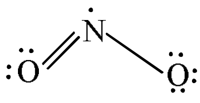


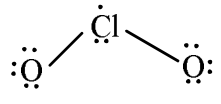
linear





or



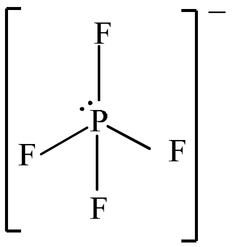


24. a. nonpolar b. polar c. polar
d. polar e. nonpolar f. non polar

25. a. N b. neither c. O
d. F e. O f. F

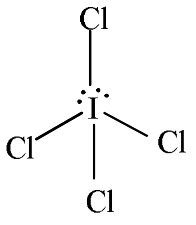
26.

a)



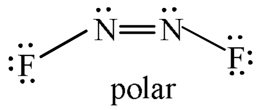
polar

b)



polar

c)



polar

27. sp^3

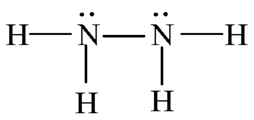
28. i σ in $CH_3 - OH$, 1σ and 1π in each of the C = O bond
1σ and 2π bonds in C ≡ O

ii. The C – O bond length in CH_3OH is greater than the C = O bond length in O = C = O, and the C – O bond length in C ≡ O is the shortest.

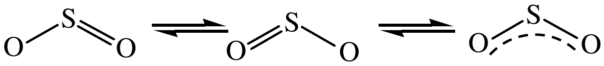
29. a. Hydrogen bond in water
b. London/dispersion forces/ increases as mass increases
c. London/dispersion forces/ in branched chain is less than that of straight chains
d. Dipole-dipole forces in SO_2

30. (B) is the most reasonable Lewis structure because it has no formal charge.

31.



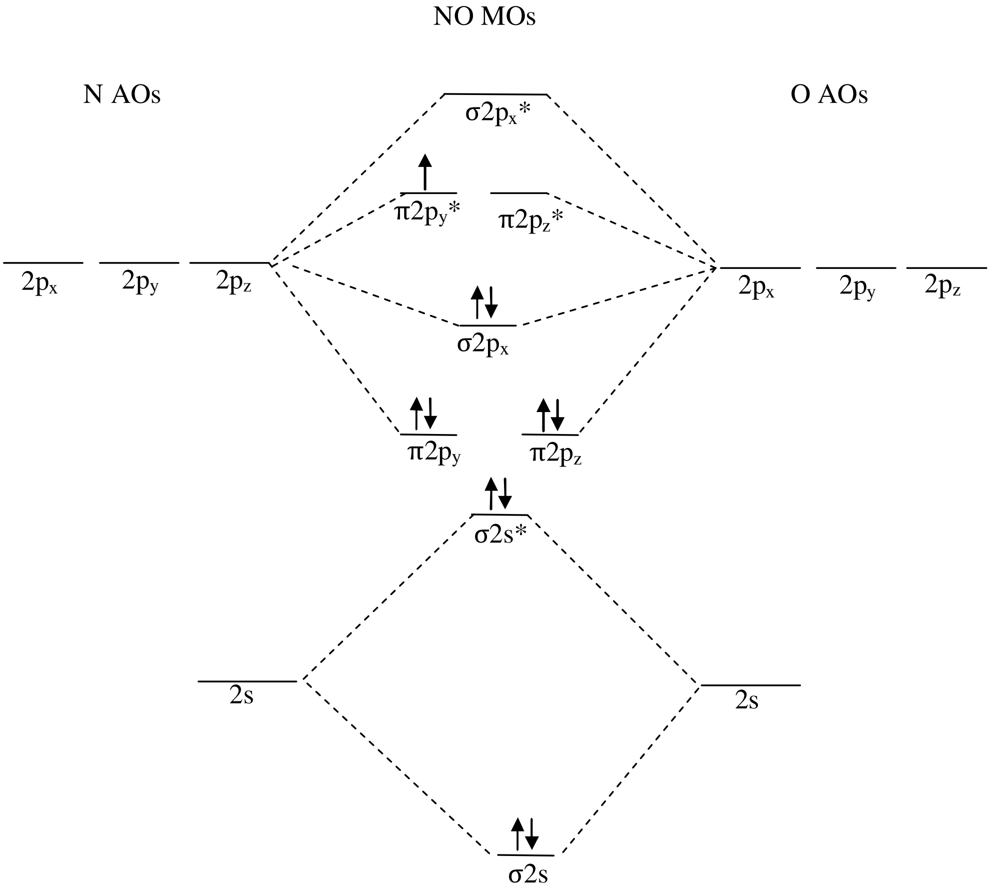
32.



33. AX_3E type – trigonal pyramid

34. a.

Count the total number of electrons in NO: 7 + 8 = 15. Of these 15 electrons, 4 are nonbonding – which leaves 11 to be distributed among the σ and π MOs. Write the number of electrons in each MO as shown in the figure below, being sure to follow the Aufbau, Pauli exclusion, and Hund principles. All of the 8 electrons in the s AOs go into the σs and σs* MOs, and the remaining 7 electrons in the p AOs are distributed in the π and π* MOs.



b. There is one electron left over after the lower-energy $\pi 2p_y$ and $\pi 2p_z$ MOs are filled with paired electrons. The single electron, which goes into the $\pi 2p_y$* antibonding MOs, is unpaired.

c.

$$\text{Bond order} = \frac{1}{2}\left[\text{Number of bonding electrons} - \text{Number of antibonding electrons}\right]$$

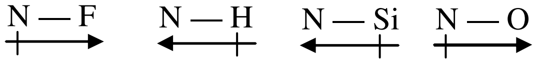
$$= \frac{1}{2}(8-3) = 2.5$$

d. paramagnetic

35.

$$O_2^+ < O_2 < O_2^-$$

36.



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UNIT 4

CHEMICAL KINETICS

UNIT OVERVIEW **Total Periods Allotted: 19 periods**

The students were introduced to the concept of rate of reaction in grade 9. This lays a good background for the students to understand the fundamental concept of *chemical kinetics* or *reaction kinetics* in this unit.

Chemical kinetics is the study of rates of chemical processes. This includes investigations of how different experimental conditions can influence the speed of a chemical reaction. The investigations yield information about the reaction's mechanism and transition states**.** The study of chemical kinetics also uses the construction of mathematical models that can describe the characteristics of a chemical reaction.

In 1864, Peter Waage and Cato Guldberg pioneered the development of chemical kinetics by formulating the law of mass action, which states that the speed of a chemical reaction is proportional to the quantity of the reacting substances.

This unit is organized into four sections. The first section (4.1), is concerned with rate of reaction, with emphasis on how to determine rate of reaction and factors that influence rate of reaction. Rate of reaction tells us how fast a quantity of a reactant or a product is changing with time. The surface area of the reactants, the presence of a catalyst, the temperature and the concentration of the substances in the reaction mixture affect the rate of a reaction.

The second section (4.2) deals with theories of reaction rates, including collision theory and transition-state theory. Collision theory qualitatively explains how chemical reactions occur and why reaction rates differ for different reactions. The rate of the reaction increases with temperature increase because a higher fraction of the collisions overcome the activation energy. Collision theory also tells us that reacting particles often collide without reacting. Certain requirements must be met if the collisions are to be effective enough to cause a reaction. In order for collisions to be successful, reacting particles must collide with sufficient energy and with a proper orientation to each other.

The third section (4.3) deals with rate equation or rate law, which includes order of reaction, rate constant, concentration-time equation and half-life of a reaction. The rate law relates the rate and concentration of the substances present in the reaction mixture. The rate is reported in terms of the rate constant of the reaction. Reactions are classified by their reaction order, the sum of the powers in which the concentrations of the reactants are raised in the rate law. The rate law must be determined experimentally and cannot be deduced from stoichiometry of the chemical equation for the reaction.

Once the expression of the concentration in terms of the time (integrated rate law) is known, it may be used to predict the concentration, at any stage of the reaction. The half-life is the time, taken for the concentration of a substance to fall to half of its initial value. The half-life of a first order reaction is independent of the initial concentration.

The fourth section (4.4) deals with reaction mechanism, which includes the molecularity of elementary reactions and the rate-determining step. A reaction mechanism is the step-by-step sequence of elementary reactions by which overall chemical change occurs. it consists of elementary reactions involving individual molecules. An elementary reaction can be classified according to the number of molecules taking part in that particular step (molecularity). A proposed mechanism must agree with the experimental rate law. The rate law implied by the mechanism is found by identifying the rate-determining step and by finding an expression for the concentration of any reaction intermediates.

Many worked examples and exercises are included in this unit in order to motivate students and so that they will have more chance to practice. Please advise your students to take time and study them closely. This unit requires a minimum of **19** periods for successful completion. We advise you to do your level best while teaching this unit.

Unit Outcomes

After completing this unit, students will be able to:

- *explain what is meant by reaction rate and perform the activities to determine it;*
- *demonstrate an understanding of the dependence of reaction rate on the nature of reactants, the surface area of solid reactants, the concentration of the reactants, the temperature of a system and catalysts;*
- *determine the reaction rates, using experimental data and calculations;*
- *describe how reaction rate theories (collision theory and transition state theory) can be used to explain changes in reaction rates;*
- *determine the rate laws and order of reactions from data on initial concentrations and reaction rates;*

- *understand that most reactions occur as a series of elementary steps in reaction mechanisms;*
- *appreciate the importance of chemical kinetics in industry and in living organisms; and*
- *demonstrate scientific enquiry skills including observing, predicting, comparing and contrasting, communicating, asking questions, drawing conclusions, applying concepts, relating cause and effect, making generalizations and problem solving.*

Main Contents

4. 1 Rate of reaction

4.2 Theories of reaction rates

4.3 Rate equation or rate law

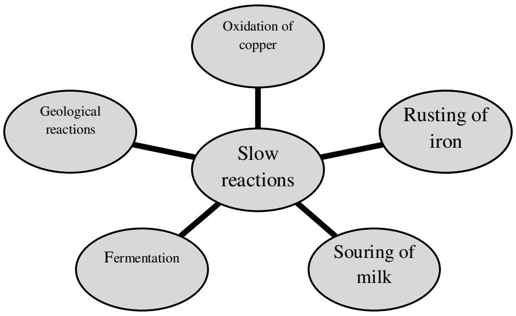
4.4 Reaction mechanism

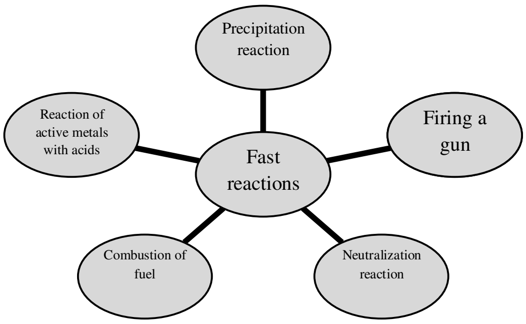
Introduction

To begin this unit, use the following activity which is not given in the student text. This activity is designed to help students review the important points about chemical reaction that they learned previously in Grade 9.

You may apply concept mapping methodology. The use of concept mapping can help students begin to understand interrelationships among concepts. This method can arrange concepts through interacting systems. In this system, all relationships between concepts are made clear. This method is suitable for developing students' cognitive structures. It can also contribute significantly towards meaningful learning in the mind of the student.

Request that the students to draw 2 columns in their notebooks, one column for fast reactions and another for slow reactions. Help each set of partners to list as many fast and slow reactions they can think of. Some suggested answers are listed in the spider diagram below.





Next, encourage the students to discuss and give their feedback on the start-up activities. The purpose of this start-up activity is to encourage students to sort out some basic facts about rate of reaction. After the discussion, the students should share their ideas with the rest of the class. Harmonize the ideas developed through discussion based on the following suggested answers.

The approximate time taken for complete burning of a piece of paper is a couple of minutes, the conversion of milk to curd takes days, ripening of grapes takes months, charring (burning) of sugar takes a few minutes, and the formation of fossil fuels takes several years.

Stabilize the discussion by mentioning that some chemical reactions proceed quickly, whereas others require days, months or even years to give products. These differences depend primarily on the differences in the chemical nature of the reacting substances. However the **rate of reaction** (speed) is not fixed for a particular type of reaction, since it depends on several factors. The rate of reaction get altered when subjected to small changes in temperature, concentration etc.

We measure the rate of reaction, by measuring the rate of change of the concentration of a reactant or of a product. The rate of reaction is defined for both reversible and irreversible reactions, as well as for both slow and fast reactions. A study of the rate of a chemical reaction informs us about the way in which the reactants combine to form products. This is explained under **reaction mechanism**.

4.1 RATE OF REACTIONS

Periods Allotted: 5 periods

Competencies

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

- *explain reversible and irreversible reactions;*
- *define dynamic chemical equilibrium;*

- *state the necessary conditions for attainment of equilibrium;*
- *describe the microscopic event that occurs when a chemical system is in equilibrium;*
- *explain characteristics of chemical equilibrium;*
- *state the law of mass action;*
- *define equilibrium constant;*
- *write the equilibrium constant expression for chemical reactions that involve concentration and partial pressure;*
- *calculate values for equilibrium constants involving concentration and partial pressure;*
- *state the relationship of* K_{eq} *to relative amounts of product and reactants in a given reaction;*
- *show the relationship between* K_c *and* K_p*;*
- *distinguish between homogeneous and heterogeneous reactions;*
- *define reaction quotient;*
- *use the equilibrium quotient to predict the direction of the reaction and the position of equilibrium;*
- *calculate equilibrium concentrations given initial concentrations;*
- *determine whether the reactants or products are favored in ahemical reaction given the equilibrium constants;*
- *list factors that affect equilibrium constants;*
- *state Le-Chatlier's principle;*
- *use Le-Chatlier's principle to explain the effect of changes in temperature, pressure, concentration and presence of catalyst on a reaction;*
- *state the effect of changes in concentration, pressure/volume and temperature on a* K_{eq}*;*
- *perform an activity to demonstrate the effect of changes in concentration and temperature on the position of equilibrium;*
- *perform an activity to determine Kc for esterification organic acids;*
- *define optimum condition; and*
- *explain how Le-Chatlier's principle is applied in the Haber process of manufacturing ammonia and contact process of manufacturing sulphuric acid.*

Forward Planning

You are expected to review the concepts of rates of reaction which the students learned in grade 9. The students have to bring their calculators every time to the class.

This topic is supported by experiments. In your plan, set appropriate times for the students to perform experimental activities. You need to do the experiments yourself before presenting them to the students. Be sure that the apparatus and chemicals required for Experiments 4.1– 4.6 are available ahead of time. A separate laboratory session may be arranged for these experiments.

You also need to set appropriate times for the other activities. You are required to help the students in forming groups, to clarify the questions in the activity where required and to facilitate their discussions among group members.

Teaching Aids

Create flipcharts for demonstrating of the dependence of reaction rate on the concentration of reactants and products. Refer to the students' textbook for the apparatus and chemicals required for performing experiment 4.1– 4.6.

Subject Matter Presentation

Active learning methodologies such as group discussion, brain storming, experiments and spider diagram are generally suggested for the delivery of the lessons in this topic.This section includes writing expressions for rate of reaction, determining rate of reaction and identifying factors that influence the rate of reaction.

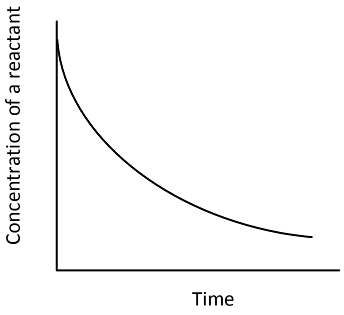
Reaction Rate and Stoichiometry

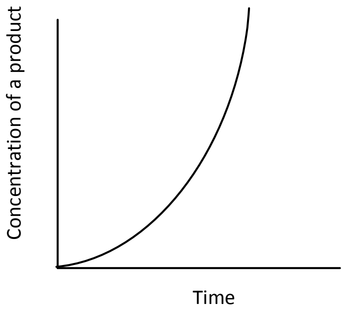
You may apply suggested active learning methodologies such as group discussion and brain storming for the delivery of the lessons in this sub-topic. Start with Activity 4.1. This activity is designed to help the students recall how the concentration of reactants and products changes with time. After the discussion, let the students share their ideas with the rest of the class and complement the ideas that they developed through discussion, based on the following suggested answers.

For the given reaction

$$2NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$$

The concentration of a reactant decreases and the concentration of a product increases as the time of reaction increases. Students can sketch a graph of changes in concentration of a reactant and a product over time where time as is written as the x-axis and concentration as the y-axis.





The students can learn the following from the graph they sketched. A reaction rate tells us how fast a quantity of a reactant or a product is changing over time. Such drawings can be used to show the rate of a reaction at a particular time. They can be obtained by calculating the slope of a curve.

Finally, take time to help students write the rate expressions for reactions indicated in Example 4.1. After you check that they have mastered that, you may give Example 4.1 as homework. Have the students complete and submit to you by the next class.

Determination of Rate of Reaction

Suggested active learning methodologies such as group discussion, experiment and brain storming may be used for the delivery of the lessons in this sub-topic. This lesson explains how to determine the reaction rate of a chemical reaction in terms of each of the reactants and products. A number of examples of chemical reactions are given. The lesson also shows how the rate of consumption of reactants is related to the rate of production of products.

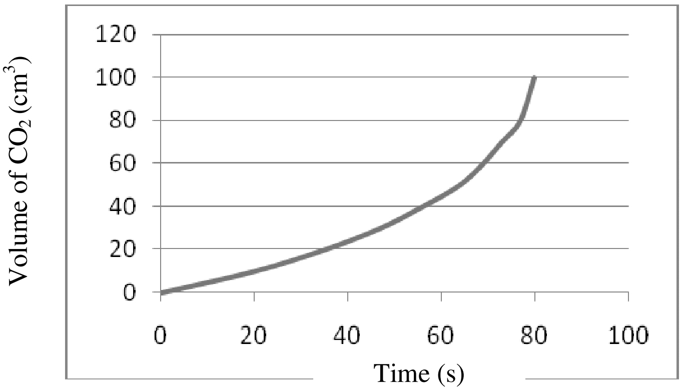
Have the students discuss Activity 4.2 which helps them to understand how the rate of a given reaction can be monitored experimentally. Let them think independently. Next, let them share their opinions with their friends in their group. Allow some of them to move to other groups to share the work of their group with the students in the other group. Finally, let some students from different groups give their answers to the whole class. Now use the following answers to help you to harmonize the ideas developed through discussion.

The results obtained during the reaction between excess of powdered calcium carbonate and hydrochloric acid are given in the following table:

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

Volume of CO_2/cm^3	0	20	35	47	56	64	69	73	77	80
Time/s	0	10	20	30	40	50	60	70	80	100

The students should plot these data as shown below. Have them use graph paper.



The reaction is most rapid between 80 s and 100 s. This can be seen from the way that the graphed curve exhibits its greatest slope (steepness) in this region. The rate of reaction after 50 seconds is calculated as

$$r = \frac{\Delta[CO_2]}{\Delta t} = \frac{(64-0)\text{cm}^3}{(50-0)\text{s}} = \frac{64}{50}\text{cm}^3\text{s}^- = 1.28\text{cm}^3\text{s}^-$$

The volume of gas formed after 15 seconds is calculated as follows:

$$r = \frac{\Delta[CO_2]}{\Delta t}$$

$$1.28\ \text{cm}^3\text{s}^- = \frac{V_2 - V_1}{(15-0)\text{s}} \Rightarrow V_2 - 0 = 19.2\ \text{cm}^3 \Rightarrow V_2 = 19.2\ \text{cm}^3$$

The time taken for 40 cm^3 of carbon dioxide to be formed is calculated as:

$$r = \frac{\Delta[CO_2]}{\Delta t}$$

$$1.28\ \text{cm}^3\text{s}^{-1} = \frac{40\ \text{cm}^3}{\Delta t} \Rightarrow \Delta t = \frac{40\ \text{cm}^3}{1.28\ \text{cm}^3\text{s}^{-1}} \Rightarrow \Delta t = 31.25\ \text{s}$$

The rate of reaction gradually decreases, since the volume of CO_2 formed is decreasing.

Next, encourage the students to practice Example 4.2 which demonstrates that the average rate of reaction depends on the time interval we choose. The students should generalize from the example that the rate of reaction at the beginning is fast and that it decreases as the reaction proceeds. Exercise 4.2 may be given to the students as homework.

Experiment 4.1 demonstrates measuring reaction rate. The purpose of this experiment is to measure the rate of reaction between marble chips ($CaCO_3$) and dilute HCl acid. A separate lab session is required. Refer to the students' textbook for the detailed procedures as well as for the apparatus and chemicals required for performing this experiment.

Apply the *predict-observe-explain methodology* as follows. In order to explain an abstract concept, a demonstration, Experiment 4.1 is often necessary to the strategy of arranging the demonstrations. First, ask students to predict the result of the demonstration. By this strategy, students are given the time and opportunity to actively think about a question. This activity in turn, gives them more motivation to study.

Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

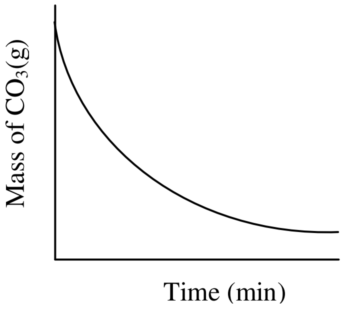
Test the experiment yourself before you let the students do it. Guide the students in setting up the apparatus in groups, as shown in the students' textbook (Figure 4.1), but without the marble chips at first. The students should strictly follow the procedure described in the textbook. Let the students submit written report to you in which they answer the questions raised under observations and analysis.

The answers to the questions listed under observation and analysis are given below.

1. Help the students to use the following table to record their observations.

Time (min)	0	1	2	3	4	5	6	7	8	9	10
Mass (g)											
Decrease in mass (g)											

Next, help them to plot a graph with to show the decrease in mass. Use time on the horizontal axis and decrease in mass on the vertical axis.



The students should draw a smooth curve through as many points as they can plot. Help them to note that, in this experiment, change in mass is proportional to change in concentration.

2. The mass of $CaCO_3$(s) decreases during the reaction. The balanced equation is

$$CaCO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$$

3. From the slope of this graph, concentration at any point can be obtained.

4. a. The graph is steepest at the start of the reaction since the reaction is fastest of them.

   b. The graph is less steep in the middle of the reaction since it is slower than at the beginning.

   c. The graph is horizontal at the end of the reaction since the mass of the $CaCO_3$ becomes nearly constant.

5. a. The average rate of the reaction is obtained from the equation

$$r = \frac{\Delta[\mathrm{CaCO_3}]}{\Delta t}$$

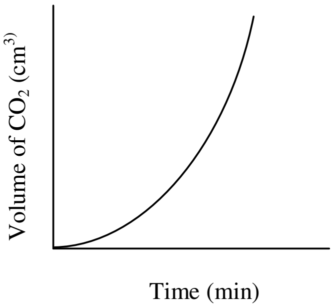
   b. The rate of reaction at 2 minutes becomes

$$r = \frac{\Delta[\mathrm{CaCO_3}]}{\Delta t} = \frac{?}{2\,\mathrm{min}}$$

6. The rate of forward reaction increases with time, and the rate of backward reaction decreases with time.

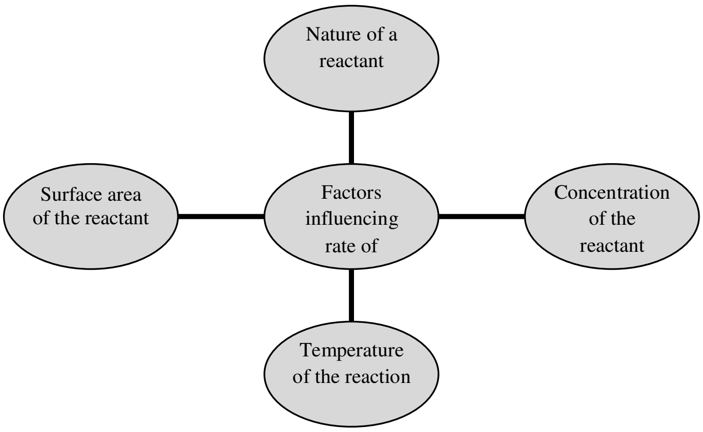
7. The rate of this reaction can be determined by measuring the amount /volume of CO_2 evolved.

8. Draw a graph of the expected result by plotting the volume of evolved CO_2 vs time.



Factors Influencing the Reaction Rates

Group discussion, spider diagram, experiment and brain storming may be used as active learning methodologies for the delivery of the lessons in this sub-topic. Use the concept-mapping methodology to help students begin to understand interrelationships among factors influencing rate of reaction. Make all relationships between concepts clear and create meaningful learning in the mind of the student. After the students try it their own way, give the feedback, referring to the spider diagram below.



After giving a short introduction to the topic, have your students perform Actvity 4.3. The purpose of this activity is to promote the scientific reasoning ability of the students. You need to monitor the work of each group.

Use the answers to this activity (given below) to help you to harmonize the ideas developed through discussion.

Sugar crystals dissolve more easily in hot water than in cold water since increasing temperature generally increases the solubility of solids in liquid. Powdered sugar dissolves easily, as compared to equivalent amounts of sugar crystals, since increasing the surface area of a solid increase its solubility in a liquid.

Effect of Nature of Reactants on Rate of Reaction

Active learning methodologies such as group discussion and experiment are suggested for the delivery of the lessons in this sub-topic. Now you may continue the presentation, assisted by Experiment 4.2, which is designed to help students understand how the nature of reactants affect the rate of reaction. Refer to the students' textbook for the detailed procedures, as well as for the list of apparatus and chemicals required in this experiment.

In order to demonstrate this experiment to the students, you may use the predict-observe-explain methodology**.** First ask students to predict the result of this demonstration. In this strategy, students are given the time and opportunity to actively think about the question, which in turn gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

Let the students do the experiment in groups of five if the required material is available. Encourage the students to discuss the questions in groups and help them to draw their conclusions from the experimental activity they performed.

The answers to the questions listed under observation and analysis are given below.

1. Of the two metals, copper does not react and magnesium reacts quickly with hydrochloric acid. This is due to the difference in the chemical nature of Cu and Mg.
2. The balanced chemical equation for the reaction of copper and magnesium with HCl are:

   $Cu(s) + \ HCl(aq) \rightarrow$ No reaction

   $Mg(s) \ + \ 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_{2(g)}\uparrow$

3. Students should learn from this experiment that the nature of the reactants is one of the factors that determines the rate of reaction.
4. In the case of aluminium, the reaction is expected to be faster than zinc since Al is a more active metal than Zn.

Effect of Surface Area of Reactants on Rate of Reaction

You may use group discussion and experiment as suggested active learning methodologies for teaching the lessons in this sub-topic. You may start this topic by asking the in-text questions from the textbook and then facilitating the discussions. Next, let the students perform Experiment 4.3, which is designed to study the effect of surface area of reactants on rate of reaction. Refer to the students' textbook for the detailed procedures as well as for the apparatus and chemicals required for performing this experiment.

Now, use predict-observe-explain strategy in order to demonstrate Experiment 4.3 to the students**.** Ask the students to predict the result of this demonstration. Give them the time and opportunity to actively think about the question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

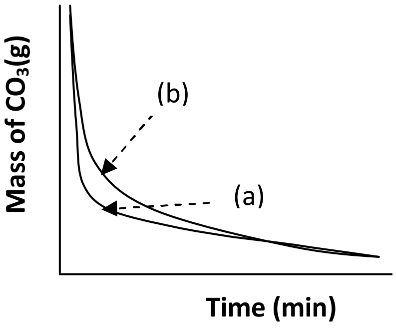
The equation representing the reaction of calcium carbonate and hydrochloric acid is

$$CaCO_3(s) + HCl\ (aq) \ \rightarrow \ CaCl_2\ (aq) + H_2CO_3(aq)$$

Encourage the students to discuss in groups the discussion question and help them to draw their conclusions from the activity they performed.

The answers to the questions listed under observation and analysis are given below.

The graph for the consumption of each size of chalk (cube chalk and powdered chalk) against time on the same graph sheet is shown below.



Graph (a) represents reaction of powdered chalk with 1 M HCl, and graph (b) represents reaction of cube chalk with 1 M HCl. The reaction rate is faster (the slope is steeper) for the reaction with powdered calcium carbonate (greater surface area).

Effect of Concentration of Reactants on Rate of Reaction

Group discussion and experiment are suggested active learning methodologies for the delivery of the lessons in this sub-topic. Similarly, you may also start this topic by asking the in-text questions from the textbook and then facilitating the discussion. This activity prepares the students to perform Experiment 4.4, which is designed for studying the effect of the concentration of hydrochloric acid on the rate of its reaction with calcium carbonate (marble chips). Refer to the students' textbook for the detailed procedures, as well as for the apparatus and chemicals required for performing this experiment.

You may apply the predict-observe-explain strategy in order to demonstrate Experiment 4.4 to the students**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

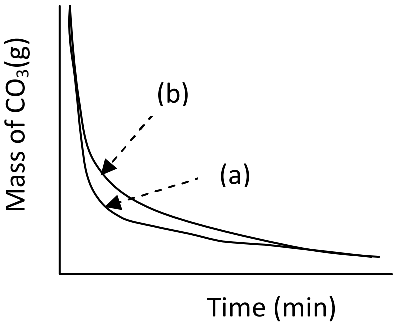
Encourage the students to discuss, in groups, the discussion question and help them to draw their own conclusions from the activity they performed.

The answers to the questions listed under observation and analysis are given below.

1. The concentrations of the reactants continually decrease as the reaction progresses. In most cases this also means the instantaneous rate of reaction decreases with time. In both cases the equation for the chemical reaction is

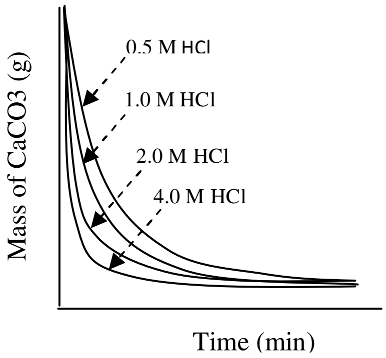
   $CaCO_3(s) + 2\ HCl\ (aq) \rightarrow CaCl_2(aq)\ +\ H_2O(g) + CO_2(g)$

In comparison with Experiment 4.1, the rate curve is as shown below. The calcium carbonate starts reacting fast and continues at a faster rate with 2 M HCl (a) than with 1 M HCl. (b) This is shown by the steepness of the graph.



Graph (a) represents reaction of calcium carbonate with 2 M HCl, and graph (b) represents reaction of calcium carbonate with 1 M HCl. The reaction rate is faster (the slope is steeper) for the reaction with calcium carbonate with 2 M HCl (greater concentration).

2. Graph (a) produces more carbon dioxide than graph (b).
3. a. When 1 M HCl is replaced with 4 M dilute HCl, the average rate of consumption of $CaCO_3$ increases, and the rate curve becomes steeper.
   b. When 1M HCl is replaced with 0.5 M dilute HCl, the average rate of consumption of $CaCO_3$ decreases, and the rate curve becomes less steep.
3. For each of 0.5 M, 1.0 M, 2.0 M and 4.0 M dilute HCl prepare against time on the same graph sheet, the results are roughly represented as shown below.



Effect of Temperature on Rate of Reaction

You may apply suggested active learning methodologies such as group discussion and experiment for the delivery of the lessons in this sub-topic. You may start this topic by asking your own activity questions and facilitating the discussions. Next, let the students perform Experiment 4.5, which is designed to help students study the effect of temperature on the rate of reaction between hydrochloric acid and sodium thiosulfate. Refer to the students' textbook for the detailed procedures, as well as the apparatus and chemicals required for performing this experiment.

Apply the predict-observe-explain strategy in order to demonstrate Experiment 4.5 to the students. Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

You can have the students write short reports after performing the experiment. You are expected to give feedback and corrections on the reports in class.

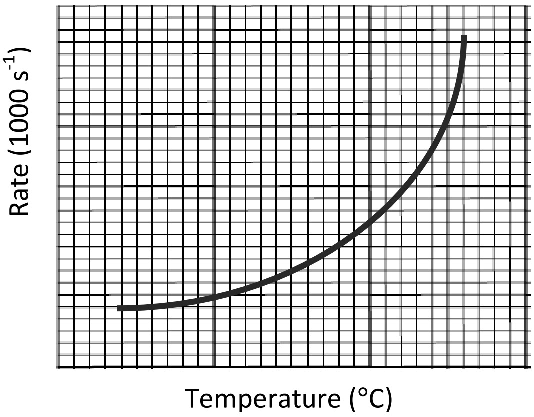
The temperature of a system is important because all chemical reaction rates increase markedly with increasing temperature. The equation for the chemical reaction is

$$Na_2S_2O_3\,(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + S(s) + SO_2(g)$$

Encourage the students to discuss, in groups, the discussion question and help them to draw their own conclusions from the activity they performed.

The answers to the questions listed under observation and analysis are given below.

1. The appearance of the mixture at the start of the reaction is colorless, and at the end of the reaction, a white creamy deposit is formed, due to the precipitation of sulfur produced.
2. For plotting temperature on the horizontal axis and reaction rate on the vertical axis, special constant unit can be used for the amount of reaction product that is needed in order to cover the cross sign. Rate is always measured as something divided by time. In many rate experiments, a gas is produced. We often measure the rate in g or cm^3 of gas produced per second (g/s or cm^3/s). That is not possible here. Instead we will measure rate as 1000/time, giving units of s^{-1}. The factor of 1000 is common in thiosulfate rate experiments, simply to give convenient numbers. The graph plotted looks like the one shown below.



3. The temperatures at which the cross takes the shortest time to disappear and the longest time to disappear are obtained from the experiment.
4. The conclusion drawn is that the reaction rate increases as the temperature of the reacting system increases.

Effect of Catalyst on Rate of Reaction

Active learning methodologies, for instance, group discussion and experiment are suggested for the delivery of the lessons in this sub-topic. Encourage the students to perform Experiment 4.6, which helps students study the effect of catalysts on the rate of reaction for the decomposition of hydrogen peroxide. Refer to the students' textbook for the detailed procedures, as well as the apparatus and chemicals required for performing this experiment. Help the students to draw their own conclusions from the activity they performed.

In order to demonstrate this experiment to the students, the best methodology that fits is the predict-observe-explain strategy**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

The answers to the questions listed under observation and analysis are given below.

1. Oxygen gas is liberated when hydrogen peroxide is decomposed. The balanced chemical equation for the reaction is written as

   $H_2O_2 \rightarrow O_2 + H_2O$

2. In the case of a catalyzed reaction, a gas syringe contains more of the gas at any one time when compared to an uncatalyzed reaction.

3. A catalyst can increase the reaction rate by forming an intermediate compound with relatively low activation energy for a reaction (homogenous catalysis), or by absorbing reacting species on its surface (heterogeneous catalysis). The appearance of the oxide before the experiment and that of the oxide obtained from the experiment after the reaction is the same.
4. We can use a data table, as shown below, to record the results.

Time (min)					
Volume of CO_2 (mL)					

The rate-of-reaction curve can be plotted, using volume of oxygen gas (mL) as the vertical axis and time (minutes) as the horizontal axis. The graphed catalyzed reaction indicates a faster rate of reaction (more steep slope) when compared to their uncatalyzed reaction.

5. The presence of the catalyst increases the rate of decomposition of hydrogen peroxide.

Finally, summarize the discussions on the rate of reaction by explaining that the rate of reaction tells us how fast a quantity of a reactant or a product is changing over time. The surface area of the reactants, the presence of a catalyst, the temperature and the concentration of the substances in the reaction mixture all affect the rate of a reaction.

Additional Notes

Chemical kinetics is the study of the rates of chemical reactions and the mechanisms by which reactions occur. A rate is the change of a property (in this case, concentration) per time unit. The rate of a chemical reaction is found by following the rate of disappearance (or decomposition) of one of the reactants or the rate of appearance (or formation) of one of the products.

Suppose we consider the reaction

$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g) \qquad (1)$$

Since three H_2 molecules react with one N_2 molecule to produce two NH_3 molecules, the rate of disappearance of H_2 will be three times the rate of disappearance of N_2, and the rate of appearance of NH_3 will be twice the rate of disappearance of H_2. Thus,

rate = rate of disappearance of N_2 = $^{1}/_{3}$ the rate of disappearance of H_2 = $^{1}/_{2}$ times the rate of appearance of NH_3

$$\text{Rate} = \left(-\frac{\Delta[N_2]}{\Delta t}\right) = \frac{1}{3}\left(-\frac{\Delta[H_2]}{\Delta t}\right) = \frac{1}{2}\left(\frac{\Delta[NH_3]}{\Delta t}\right) \qquad (2)$$

Note that rates are positive numbers (time doesn't go backwards!). That is the reason for the negative sign in the expressions for the rate of disappearance of N_2 and H_2 in the Equation (2). For example, since $\Delta[N_2]$ is a negative number, we need to multiply $(\Delta[N_2]/\Delta t)$ by (-1) to obtain a positive rate.

In general, for the chemical equation

$a\text{A} + b\text{B} \rightarrow c\text{C} + d\text{D}$

the *rate* is given by

$$\text{rate} = \frac{1}{a}\left(-\frac{\Delta A}{\Delta t}\right) = \frac{1}{b}\left(-\frac{\Delta B}{\Delta t}\right) = \frac{1}{c}\left(\frac{\Delta C}{\Delta t}\right) = \frac{1}{d}\left(\frac{\Delta D}{\Delta t}\right) \qquad (2)$$

Since the rates of appearance and disappearance of all reactants and products are related by the stoichiometry equation, it doesn't matter which rate we actually measure – experimental convenience governs our choice. However, since the rates differ by stoichiometric ratios, we must specify the **substance** for which our *rate* is defined.

Assessment

Assess each student's work to determine whether or not the students have achieved minimum required level. Don't forget to record the students' performance in participation, in discussion and in answering questions. Help students working below the minimum requirement level by giving them extra activities, listed as additional exercises, so that they will catch up with rest of the class. Give them extra attention in class and additional lesson time during breaks or at the end of the day.

Don't forget to appreciate students working above minimum requirement level. You are required to assist those working below the minimum requirement level, and you may give them the extra exercises listed as additional exercises. Exercises 4.1–4.5 can be used to assess the students' work and understanding.

Additional Questions

*1. Give three ways of increasing the rate of a reaction.

*2. How can the rate of a reaction be measured?

*3. Write the equation for the reaction between calcium carbonate and HCl. Explain how the rate of this reaction is measured.

*4. Write the equation for the reaction between sodium thiosulfate and HCl. How is the rate of formation of sulfur measured?

*5. Write the equation for the decomposition of hydrogen peroxide. What catalyst is used?

*6. What happens if you increase the temperature?

*7. What happens if you increase the concentration? Explain, using the reaction between sodium thiosulfate solution and dilute hydrochloric acid.

*8. What happens to the rate of a reaction for gaseous reactants if you increase the Pressure?

*9. What happens if you use a powder in the reaction between calcium carbonate and dilute hydrochloric acid?

*10. Express the rate of the reaction:

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

a. In terms of $[H_2]$ b. In terms of [HI]

11. Given the reaction

$$2NOC\ (g) \rightarrow 2NO(g) + Cl_2(g)$$

express the rate of reaction in terms of the concentration of NOCl and relate this to the rate of formation of each of the products.

12. In the reaction of nitrogen(II) oxide with hydrogen at 100 K

$$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

The rate of disappearance of NO is

$$r = -[NO]/\Delta t = 5.0 \text{ x } 10^{-5} \text{ mol L}^{-1}\text{s}^{-1}.$$

What is the rate of formation of N_2?

Answers to Additional Questions

*1. There are five ways to increase the rate of a chemical reaction. They are all understood in terms of collision theory. The rate of a chemical reaction may be increased by

a. Raising the temperature.

b. Increasing the concentration (in solution).

c. Increasing the pressure (in gases).

d. Increasing the surface area in solid(s).

e. Using a catalyst.

The opposite of a, b, c and d will decrease the rate of a reaction. A catalyst (strictly speaking) will change the rate of a reaction, making it go faster or slower. In practice, a catalyst is mainly used to make a reaction go faster.

*2. The rate of a reaction may be measured by following the loss of a reactant, or the formation of a product.Three of the reactions which may be studied to show how the rate can be changed are shown below.They are

a. The reaction between calcium carbonate and dilute hydrochloric acid.

b. The reaction between sodium thiosulfate solution and hydrochloric acid.

c. The decomposition of hydrogen peroxide solution.

*3. The equation for the reaction between calcium carbonate and HCl is

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

The rate of this reaction can be measured by following the rate at which carbon dioxide is formed. This can be done by conducting the reaction in an open flask on an electric balance (weighing machine).

As the carbon dioxide escapes to the air, the mass of the flask and its contents decreases. You can take a reading from the balance every 30 seconds, then plot a graph of loss-of-mass against time.

The gradient of the plot (the steepness of the slope) shows the rate of the reaction (how fast it is going).

*4. The equation for the reaction between sodium thiosulfate and HCl is

$$HCl(aq) + Na_2S_2O_3(aq) \rightarrow NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

The solid sulfur (S(s)) formed in this reaction makes the colourless solution go cloudy. The reaction is usually carried out in a flask placed on a piece of white paper which has a black cross on it. At the beginning of the reaction, the cross can easily be seen. As the flask becomes more and more cloudy, the cross gets harder to see.

You can measure the time from the start of the reaction until the cross can no longer be seen. This is a way of measuring the rate of formation of sulfur.

*5. The equation for the decomposition of hydrogen peroxide is

$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$$

The reaction is carried out in a closed flask which has a gas syringe connected to the top of it. The reaction is started by adding a catalyst to the hydrogen peroxide.

The catalyst is manganese (IV) oxide, MnO_2. The volume of oxygen in the syringe increases as the reaction proceeds. The volume of oxygen can be noted every 30 seconds, and a graph of volume against time can be plotted. The gradient of the plot (the steepness of the slope) shows how fast the reaction is going.

*6. Raising the temperature has the same effect on all three reactions. Raising the temperature makes the particles move faster. This means that the particles collide more frequently with each other and the rate of the reaction increases.

Also, the faster the particles are traveling, the greater is the proportion of them which will have the required activation energy for the reaction to occur.

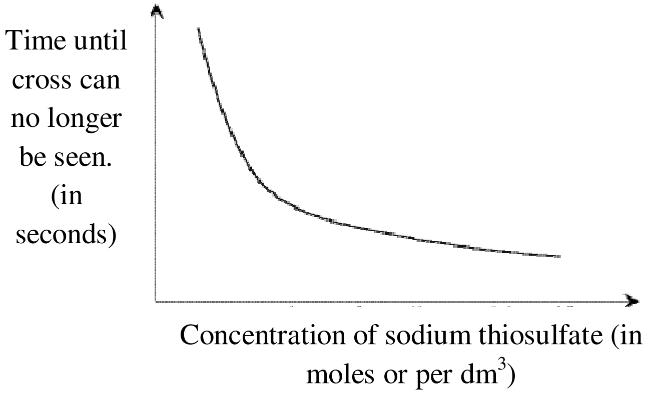
As a general guide, raising the temperature of a reaction by 10 °C will double the rate of a reaction. The gradient of the plot will be twice as steep.

*7. Increasing the concentration of a substance in a solution means that there will be more particles per volume of that substance. The more particles that are in the same volume of solution, the closer these particles are to each other. This means that the particles collide more frequently with each other, and therefore the rate of the reaction increases.

In the reaction between sodium thiosulfate solution and dilute hydrochloric acid

$HCl(aq) + Na_2S_2O_3(aq) \rightarrow NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$

solid sulfur (S(s)) is formed in the flask. Increasing the concentration of sodium thiosulfate causes the solid sulfur to be produced more quickly. Therefore less time passes before the cross can no longer be seen.



When making different concentrations of sodium thiosulfate, take care to use the same total volume of sodium thiosulfate plus hydrochloric acid for the comparison to be meaningful.

*8. Increasing the pressure of a reaction system, where the reactant is a gas, is similar to increasing the concentration of a reactant in a solution. The gas particles (usually molecules) are closer together under the increased pressure. This causes the particles to collide more frequently with each other, and therefore the rate of the reaction increases.

If the reaction is reversible, then increasing the pressure favors the side of the reaction which has the smaller volume (see the Haber process).

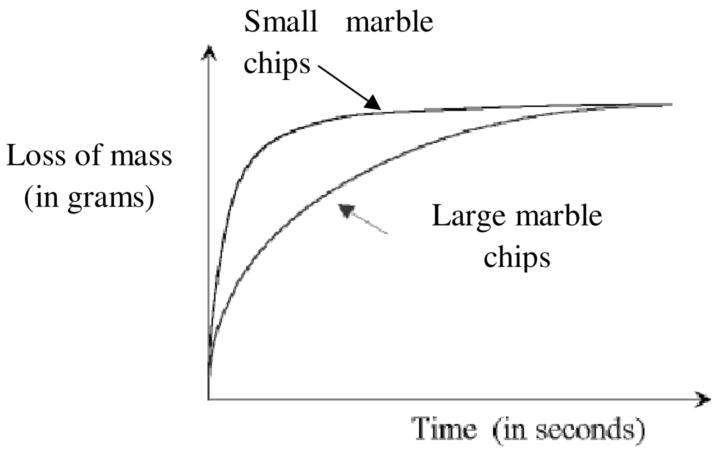
*9. A solid in a solution can only react when particles collide with its surface. The bigger the area of the solid surface, the more particles can collide with it per second,

and the faster the reaction rate is. You can increase the surface area of a solid by breaking it up into smaller pieces (see also nanoparticles). A powder has the largest surface area and therefore has the fastest reaction rate. This is why catalysts are often used in powdered form.

In the reaction between calcium carbonate and dilute hydrochloric acid

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

calcium carbonate may be used in the form of marble chips. Reaction rates can be compared by using equal masses of large marble chips and small marble chips. The reaction can be followed by plotting the loss of mass against time.



The reaction rate is faster (the slope is steeper) for the reaction with small marble chips (greater surface area). Note that the final loss of mass is the same for both reactions. This is because the same masses of calcium carbonate (marble chips) will give the same mass of carbon dioxide, whether the chips are large or small. The smaller chips will just do it more quickly.

*10. The rate of the reaction:

$2HI(g) \rightarrow H_2(g) + I_2(g)$

In terms of $[H_2]$ is $r = \Delta [H_2] / \Delta t$

In terms of [HI] $r = -\frac{1}{2} \Delta [HI] / \Delta t$

11. For the reaction:

$2NOC(g) \rightarrow 2NO(g) + Cl_2(g)$

the rate of reaction, in terms of the concentration of NOCl $r = -\Delta [NOCl] / \Delta t$.

The rate of formation of each of the products:

$r = -\Delta [NOCl] / \Delta t = \Delta [NO] / \Delta t = 2 \Delta [Cl_2] / \Delta t$ or

$r = -\frac{1}{2} \Delta [NOCl] / \Delta t = \frac{1}{2} \Delta [NO] / \Delta t = \Delta [Cl_2] / \Delta t$

12. In the reaction of nitrogen(II) oxide with hydrogen at 100 K,

$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$

If the rate of disappearance of NO is $r = -[NO]/\Delta t = 5.0 \times 10^{-5}$ mol.L^{-1}.s^{-1}, then the rate of formation of N_2 is

$r = \Delta [N_2] / \Delta t = -\frac{1}{2} [NO]/\Delta t = \frac{1}{2} (5.0 \times 10^{-5} \text{ mol.L}^{-1}.\text{s}^{-1}) = 2.5 \times 10^{-5} \text{ mol.L}^{-1}.\text{s}^{-1}$

Answers to Exercise 4.1

1. Rate of reaction is the change in concentration of a reactant or a product over a given period of time.

2. For the reaction

$2HI(g) \rightarrow H_2(g) + I_2(g)$

a. rate of reaction in terms of the concentration of H_2 is $r = \frac{\Delta[H_2]}{\Delta t}$

b. rate of reaction in terms of the concentration of, HI is $r = -\frac{1}{2}\frac{\Delta[HI]}{\Delta t}$

3. For the reaction

$2NOCl(g) \rightarrow 2NO(g) + Cl_2(g)$

a. the rate of reaction in terms of the concentration of NOCl is

$$r = -\frac{1}{2}\frac{\Delta[NOCl]}{\Delta t}$$

b. the rate of reaction of NO is

$r = \frac{1}{2}\frac{\Delta[NO]}{\Delta t}$ and, in terms of Cl_2, is $r = \frac{\Delta[Cl_2]}{\Delta t}$

4. The reaction for the formation of ammonia is given as:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

a. the rate reaction for the formation of NH is

$$r = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$$

b. the disappearance of N_2 is

$r = -\frac{\Delta[N_2]}{\Delta t}$ and H_2 is $r = -\frac{1}{3}\frac{\Delta[H_2]}{\Delta t}$.

Answers to Exercise 4.2

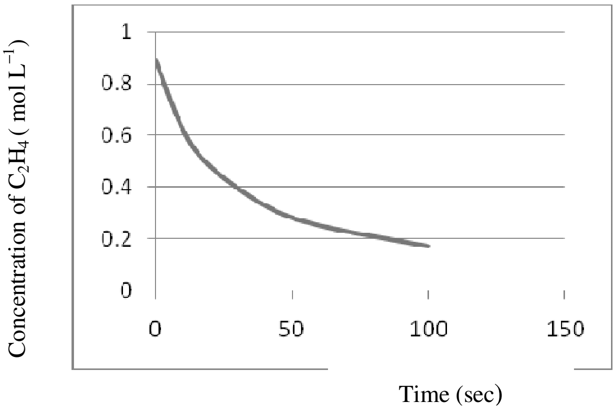
A rate of reaction graph for the consumption of ethylene, C_2H_4, in the reaction

$$2C_2H_4(g) \rightarrow C_4H_8(g)$$

from experimental data given in the table below can be plotted as follows. Use time as the x-axis and $[C_2H_4]$, mol L^{-1} as the y-axis.

Time, s	0	10	20	40	60	100
$[C_2H_4]$, mol L^{-1}	0.884	0.621	0.479	0.328	0.250	0.169

The graph looks as shown below.



Answers to Exercise 4.3

Using the data in Table 4.2 for the reaction

$N_2O_5(g) \rightarrow 2NO_2(g) + ½ O_2(g)$ to calculate the rate of

a. formation of NO_2 and O_2 over the first 10 minutes (min).

$$r = -\frac{1}{2}\frac{\Delta[NO_2]}{\Delta t} = -\frac{1}{2}\times\frac{(1.24\times10^{-2} - 0.92\times10^{-2})}{(10-0)s}$$

$$r = -\frac{1}{2}\times\frac{(-0.32\times10^{-2})M}{10s} = -\frac{1}{2}\times -0.32\times10^{-3}Ms^{-1} = 1.6\times10^{-4}Ms^{-1}$$

b. decomposition of N_2O_5 during the time interval t = 30 min to t = 50 min.

$$r = \frac{\Delta[\mathrm{N_2O_5}]}{\Delta t} = \frac{(1.92 \times 10^{-2} - 1.48 \times 10^{-2})M}{(50-30)s}$$

$$= \frac{0.44\times 10^{-2}M}{20\mathrm{s}} = 2.2\times10^{-4}Ms^{-1}$$

Answers to Exercise 4.4

1 The five factors that affect reaction rates:

- ✓ nature of the reactants
- ✓ state of subdivision of solid reactants
- ✓ temperature
- ✓ concentration of the reactants in solution
- ✓ presence of a catalyst

2 A reaction of substances when both are in the gaseous or liquid phase is referred to as a *homogeneous reaction*. A *heterogeneous reaction* is reaction between substances that are in different physical phases (solid, liquid, gas)

Answers to Exercise 4.5

1. When temperature increases, overall kinetic energy of all particles of reactants increases, so more molecules will have energy greater than activation energy. So, the overall frequency of collisions increases, therefore frequency of effective collisions increases, and thus the reaction rate increases.

2. A catalyst is a substance used in the reaction to decrease the value of activation energy and, thereby, to provide a pathway to increase the rate of reaction by increasing the number of effective collisions between reactants. The catalyst is not consumed.

3. When the size of the substance subdivisions decreases, surface area increases. It makes the effective collisions increase and therefore the rate of reaction also increases.

4. The effect of surface area can be studied by comparing the reactivity of zinc dust, zinc granules, or a piece of zinc sheet with HCl. The effect of temperature can be observed by comparing the reaction in an ice bath and in a water bath, using a constant reaction-system temperature that is higher than the temperatures of both baths.

4.2 THEORIES OF REACTION RATES

Period Allotted: 4 periods

Competencies

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

- *state collision theory;*
- *define activation energy;*
- *describe how collision theory can be used to explain changes in reaction rate;*
- *state transition state theory;*
- *define activated complex (transition state);*
- *describe how transition state theory can be used to explain changes in reaction rate;*
- *sketch and label the energy profiles of reactions which are exothermic and endothermic.*

Forward Planning

Take enough time to read all of the contents of this topic mentioned in the textbook in order to be able to help the students whenever necessary. After reading the student text book you can plan a method for delivering the lesson. Be prepared to discuss the collision theory and the transition state theory. Think of a simple activity that provides an analogy of the two theories. You may also need to prepare the energy diagrams for presentation of this topic.

Teaching Aids

No experiment is suggested for this section. Reaction energy diagrams prepared on flip charts are useful aids for demonstrating the potential energy curves for exothermic and endothermic reactions.

Subject Matter Presentation

You may apply suggested active learning methodologies such as brain storming, enquiry, group discussion and debate for the delivery of the lessons in this topic. This section includes collision theory (with respect to concentration, temperature), and transition state theory. The collision theory qualitatively explains how chemical reactions occur and why reaction rates differ for different reactions. It also tells us that certain requirements must be met if the collisions are to be effective enough to cause a reaction. In order for collisions to be successful, reacting particles must collide with sufficient energy and with the proper orientation to each other. Transition state theory is another supporting view that describes a chemical reaction.

Collision Theory

Suggested active learning methodologies such as brain storming, enquiry, group discussion and debate may be used for the delivery of the lessons in this sub-topic. Introduce the topic and let your students perform Activity 4.4. Its purpose is to help students understand the importance of effective collision in a chemical reaction. Let them focus on generating questions in the activity. Listen to the students' conclusions and continue your explanation by harmonizing their opinions with the facts presented in their text, use the following suggested answers.

Ethanol can easily burn in air, due to the following reaction.

$$CH_3CH_2OH + O_2 \rightarrow CO_2 + H_2O$$

In actual practice, we are required to use a lighter (igniter) to start the reaction. This is because we need a minimum amount of energy to start combustion of ethanol. Consider the following reaction

$$NO_2Cl + Cl \rightarrow NO_2 + Cl_2$$

Orientation (b) is effective for a collision that leads to the formation of the product. In this effective orientation, the two Cl atoms are directly facing each other. The students should be able to discuss concentration and the collision theory as stated in their textbook.

Use problem-based learning methodology and begin the teaching-learning with an ill-process structured and open-ended problem. In this way, guide the students to know why they're learning. Problem-based learning places the emphasis on understanding what knowledge is needed and on how to use it in a given situation.

Concentration and collision theory

Suggested active learning methodologies such as brain storming, enquiry, group discussion and debate may be used for the delivery of the lessons in this sub-topic. Students should know that collision theory views the reaction rate as the result of particles colliding with a certain frequency and minimum energy. The basis for collision theory is the notion that molecules must collide in order to react. We can predict the rate law if we know what collisions take place during a reaction. Therefore, the factors that increase collision frequency also increase the rate of a reaction.

As a corollary to this, students should know that, if there is no collision at all, then there is no reaction at all, and its reaction rate is zero. Out of the factors that affect rates mentioned above, concentration and temperature have direct relevance in this theory.

Ensure that your students understand how collision frequency increases if concentration is increased.

Temperature and collision theory

Suggested active learning methodologies such as brain storming, enquiry, group discussion and debate may be used for the delivery of the lessons in this sub-topic. Let the students be aware of the fact that, based on physics, higher temperature implies higher kinetic energy. Remind the students that $K.E = ½ mv^2$, which implies the molecules will collide frequently and effectively which, in turn, implies a higher reaction rate. The minimum kinetic energy required between colliding molecules in order to produce an effective collision is called the activation energy, E_a.

Here the key question "What are the preconditions that must be fulfilled for a reaction to occur?" can be used to involve the students in discussion. The students should be able to summarize the postulates of collision theory as stated in their textbook. Help, guide and encourage the students to:

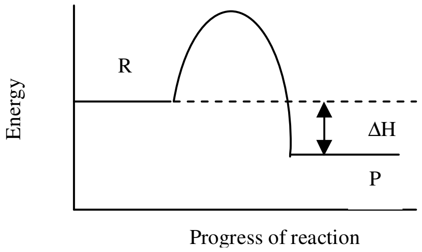
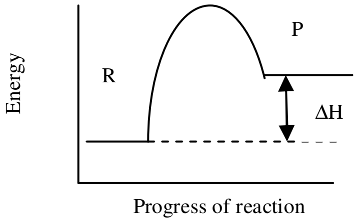
- discuss the probability of particles colliding with each other in a confined space.
- point out that a collision will not necessarily result in a reaction.
- describe, with examples, the importance of the orientation of particles for achieving an effective collision.
- use simple diagrams to explain the concept of activation energy.

Emphasize that temperature plays a crucial role in determining the rate of any reaction.

Transition State Theory - Activated Complex and Energy Barrier

Suggested active learning methodologies such as brain storming, enquiry, group discussion and debate may be used for the delivery of the lessons in this sub-topic. Allow the students to form a group and use *Activity 4.5* to involve them in discussion. You are required to clarify the questions in the activity where required and to facilitate discussion among group members. Appropriate time should also be set for the activity.

Let them think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. The following answers can help you to harmonize the ideas they develop through discussion.

Differences	**Exothermic reaction**	**Endothermic reaction**
Energy	release energy	absorb energy
Energy profile diagram		
Examples	(a) burning of candle (b) respiration	(a) dissolution of ammonium (b) chloride in water (c) photosynthesis

Next, motivate and guide the students to:

- ✓ review collision theory from their previous lesson by means of an analogy.
- ✓ discuss the energy changes that occur during the course of a reaction.
- ✓ discuss the formation and energy of the activated complex.

Transition State and Enthalpy of Reaction

Suggested active learning methodologies such as brain storming, enquiry, group discussion and debate may be used for the delivery of the lessons in this sub-topic. Here you may involve your students in a simple activity in groups. Let students try to provide some practical examples from common experience to describe the concept of sufficient activation energy to overcome a given energy barrier. They may think of rolling a ball up an inclined plane. Sufficient energy supply is required to lift up the ball over the energy barrier. A given stone or block of wood needs an initial push with sufficient energy to roll it down hill.

Such examples help them to visualize the reaction energy diagrams. Let students relate such phenomena with a chemical reactant that is to be transformed to product by passing through an energy barrier. You will harmonize the ideas developed by students in your next presentation. Here, you may use diagrams to support explanation.

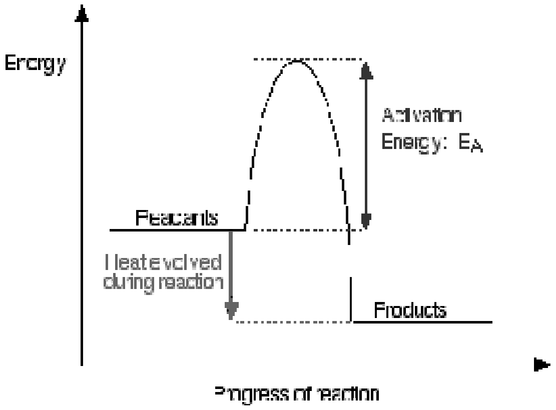
Encourage the students to summarize that there are two theories running parallel that attempt to explain rates of reaction. These are collision theory and transition state theory. These theories are distinct but in complete agreement, each emphasizing different aspects of a reaction process.

Additional Notes

Collision theory qualitatively explains how chemical reactions occur and why reaction rates differ for different reactions. This theory is based on the idea that reactant particles must collide, but that only a certain fraction of the total collisions have the energy to connect effectively and cause the reactants to transform into the products. This is because only a portion of the molecules have enough energy and the right orientation (or "angle") at the moment of impact to break any existing bonds and form new ones. The minimal amount of energy needed for this to occur is known as *activation energy.*

We can show this on an *energy profile* for the reaction.

For a simple over-all exothermic reaction, the energy profile looks like this:



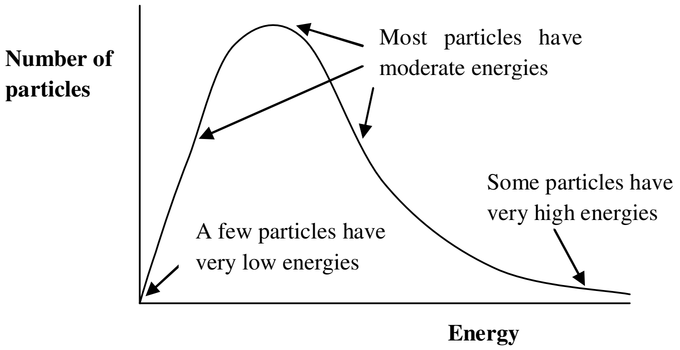
If the particles collide with less energy than the activation energy, nothing important happens. The particles bounce apart. We can think of the activation energy as a barrier to the reaction. Only those collisions which have energies equal to or greater than the activation energy result in a reaction.

Any chemical reaction results in the breaking of some bonds (needing energy) and the making of new ones (releasing energy). Obviously, some bonds have to be broken before new ones can be made. Activation energy is involved in breaking some of the original bonds. Where collisions are relatively gentle, there isn't enough energy available to start the bond-breaking process, and so the particles don't react.

The Maxwell-Boltzmann Distribution

Because of the key role of activation energy in determining whether a collision will result in a reaction, it would obviously be useful to know what sort of proportion of the particles present have high enough energies to react when they collide. In any system, the particles present will have a very wide range of energies. For gases, this can be shown on a graph called the *Maxwell-Boltzmann Distribution* which is a plot of the number of particles each having particular energy.

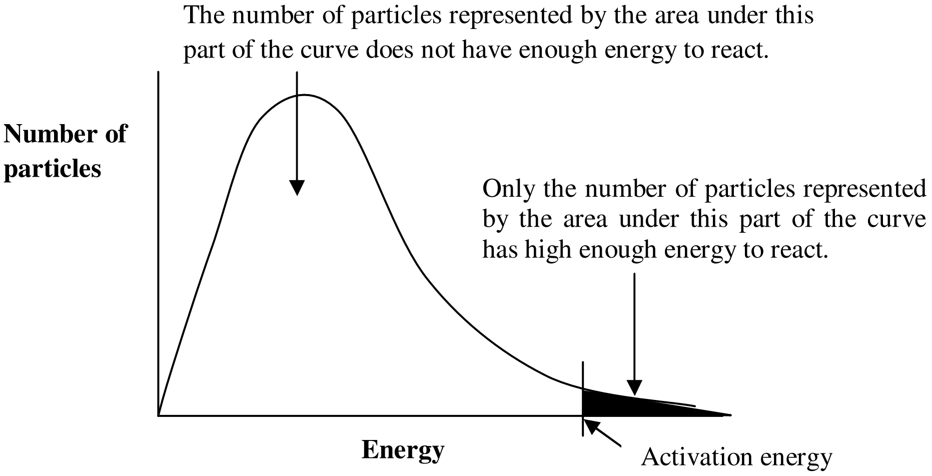
Note: The graph only applies to gases, but the conclusions that we can draw from it can also be applied to reactions involving liquids.



The area under the curve is a measure of the total number of particles present.

The Maxwell-Boltzmann Distribution and activation energy

Remember that, for a reaction to happen, particles must collide with energies equal to or greater than the activation energy for the reaction. We can mark the activation energy on the Maxwell-Boltzmann distribution:



Notice that the large majority of the particles don't have enough energy to react when they collide. To enable them to react, we either have to change the shape of the curve or move the activation energy further to the left. Note that we can change the shape of the curve by changing the temperature of the reaction. We can change the position of the activation energy by adding a catalyst to the reaction.

Assessment

Continuously assess the participation of students in each activity in this unit. Ask simple assessment questions at the beginning, the middle and the end of each lesson.

The students should be able to: briefly describe and explain both aspects of the collision theory; explain the effect which the orientation of the colliding particles has on the effectiveness or otherwise of the collision; describe the effect which the temperature of the reaction system has on the effectiveness of the collisions; explain what is meant by 'activation energy' and illustrate it by means of a sketch graph (you may use Figure 4.7 for this purpose); define activation energy, activated complex and heat of reaction; and sketch and label the energy profiles of reactions which are exothermic and endothermic.

Additional Questions

*1. What does collision theory state?

*2. What is the minimum amount of energy needed for a reaction called?

*3. Explain why all reactions have activation energy using your knowledge of collision theory.

*4. Describe how the activation energy, of a reaction affects the overall rate of the chemical reaction.

5. The rate of a chemical reaction can be doubled by increasing the reaction temperature by ten degrees Celsius. Explain this drastic increase in reaction rate, using your knowledge of collision theory.

6. It has been observed that more gas station fires occur on hot days than on cold days. Explain this phenomenon, using your knowledge of collision theory (Hint: It’s not just the temperature increase that causes this!).

7. It has been observed, with one variety of paint, that the rate of paint drying can be drastically increased by adding a small amount of “accelerant”. Based on what you know of catalysts, is it reasonable to think of this accelerant as being a catalyst? Explain.

Answers to Additional Questions

*1. Collision theory says that a chemical reaction can only occur between particles when they collide (hit each other). Particles may be atoms, ions or molecules.

*2. There is a minimum amount of energy which colliding particles need in order to react with each other. If the colliding particles have less than this minimum energy

then they just bounce off each other and no reaction occurs. This minimum energy, is called the activation energy.

The faster the particles are going, the more energy they have. Fast moving particles are more likely to react when they collide. You can make particles move more quickly by heating them up (raising the temperature).

*3. All reactions have activation energy because energy is required to make the reactants combine in a way that will cause the reaction. No chemical process can take place without having at least a little energy to get things started.

*4. The smaller the activation energy, the faster is the reaction.

5. For many chemical reactions, the amount of energy available at room temperature is only slightly less than the activation energy for the reaction. By adding only a small amount of energy, you can get a huge increase in reaction rate because you add that last little bit of energy required to make the reaction proceed.

6. The increased temperature itself plays a very small role – there is enough energy at even freezing temperatures for gasoline to burn easily. The reason more gas station fires occur at high temperatures is that the gasoline evaporates, and the vapor tends to explode. According to collision theory, this corresponds to an increase in concentration.

7. No, because the accelerant is used up during the reaction.

Answers to Exercise 4.6

1. Activation energy is the minimum amount of energy required to start a chemical reaction.
2. Transition-state theory, which is also referred to as activated-complex theory, states that the collision between two reacting species results in the formation of an activated complex or transition state.
3. A catalyst increases the rate of a reaction by lowering the activation energy.
4. During the course of a reaction, one activated complex can be formed for a given simple reaction.
5. There is a direct relationship between collision and rate of chemical reaction.

Answers to Exercise 4.7

1. Given that E_a for the hydrolysis of sucrose is 108 x 10^3 kJ /mol, compare the rate constant of this reaction at 37 °C(T_1) with the rate constant of the same reaction at 27 °C(T_2).

Solution:

$T_1 = 37 + 273 = 310$ K $T_2 = 27 + 273 = 300$ K

R = 8.31 J/K.mol

$E_a = 108 \times 10^3 = 1.08 \times 10^5$ kJ/mol

A = not given

$$\ln k_1 = \ln a - \frac{E_a}{RT_1} \text{ and } \ln k_2 = \ln a - \frac{E_a}{RT_2}$$

$$\ln k_1 - \ln k_2 = \left(\ln a - \frac{E_a}{RT_1} \right) - \left(\ln a - \frac{E_a}{RT_2} \right)$$

$$\ln k_1 - \ln k_2 = \ln a - \frac{E_a}{RT_1} - \ln a + \frac{E_a}{RT_2} = -\frac{E_a}{RT_1} + \frac{E_a}{RT_2}$$

$$\ln \frac{k_1}{k_2} = -\frac{E_a}{RT_1} + \frac{E_a}{RT_2} \quad (\text{ since } \ln k_1 - \ln k_2 = \ln k_1 / k_2)$$

$$\ln \frac{k_1}{k_2} = -\frac{1.08 \text{ x } 10^5}{8.314 \text{ x } 310} + \frac{1.08 \text{ x } 10^5}{8.314 \text{ x } 300} = 1.396$$

$$\frac{k_1}{k_2} = 4.04$$

The reaction at 37 °C is about 4 times faster than at 27 °C.

2. We now want to know how much faster the same reaction would be at 47°C than at 37°C.

Solution:

$T_1 = 47 + 273 = 320$ K $T_2 = 37 + 273 = 310$ K
R = 8.31 J/k.mol
$E_a = 108 \text{ x } 10^3 = 1.08 \text{ x } 10^5$ kJ /mol
A = not given.

$$\ln \frac{k_1}{k_2} = -\frac{E_a}{RT_1} + \frac{E_a}{RT_2}$$

$$\ln \frac{k_1}{k_2} = -\frac{1.08 \text{ x } 10^5}{8.314 \text{ x } 320} + \frac{1.08 \text{ x } 10^5}{8.314 \text{ x } 310} = 1.31$$

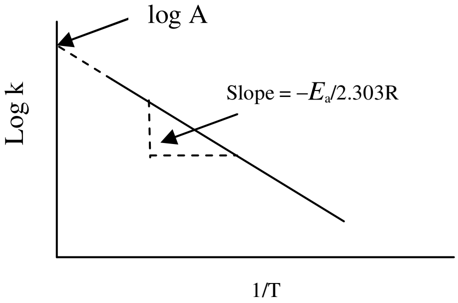
$$\frac{k_1}{k_2} = 3.71$$

The reaction at 47 °C is almost 4 times faster than at 37 °C.

3. The equation for a straight line that relates k with T is given by:

$$\log k = \left(\frac{-E_a}{2.303R} \right) \frac{1}{T} + \log A$$

The value of E_a can be found from the slope of a plot of log k versus $1/T$, $E_a = 2.303\ R \times \text{slope}$.



4.3 RATE EQUATION OR RATE LAW

Periods Allotted: 7 periods

Competencies

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

- *define rate law, order of reaction and rate constant;*
- *determine reaction order;*
- *calculate rate constants from a given experimental data;*
- *explain the zero, first and second order reactions using concentration versus time curve;*
- *calculate that concentration and time the reaction mixture for different order reactions and plot their graphs; define half life; and*
- *calculate the half lives of zero order, first order and second order reactions from experimental data.*

Forward Planning

Make the necessary preparations before the class. Read the appropriate section of the students' text book and other supplementary reference materials available to you. Be prepared to explain what is meant by rate expression or rate law, reaction order and the rate constant. Practice the calculation of rates and rate constants for zero order and first-order reactions by using the method of initial rates. Focus especially on the fact that the rate law is determined by experimental means.

Assigning students the task of reading the contents of this topic ahead for time is helpful to facilitating discussions and preparing them for your lecture in the class. Diagrams prepared on flip-charts may help you.

Also, follow these guidelines to prepare yourself:

- review the derivation of the formula.
- work through the examples given in the student textbook.
- become familiar with the definition of half-life and with the process for deriving an equation in terms of the half-life and the reaction rate constant
- the concept of half-life is the time taken for the initial concentration of a reactant to drop by half.
- the half-life equation enables us to calculate rate constants, or to discover how long it would take for a specified percentage of the reaction to complete.

Teaching Aids

There is an experiment suggested for this section. Prepare flipcharts for your demonstration of the change of concentration of reactants with time for the zero-order, first-order and second-order reactions.

Subject Matter Presentation

Active learning methodologies such as group discussion, brain storming, gaped lectures and problem solving are generally suggested for the delivery of the lessons in this topic. This section deals with rate equation or rate law, which includes order of reaction, rate constant, concentration-time equation and half-life of a reaction. The rate law relates the rate and concentration of the substances present in the reaction mixture. The rate is reported in terms of the rate constant of the reaction.

Reactions are classified by their reaction order, the sum of the powers in which the concentrations of the reactants are raised in the rate law. The rate law must be determined experimentally and cannot be deduced from stoichiometry of the chemical equation for the reaction. Once the expression of the concentration in terms of the time (integrated rate law) is known, it can be used to predict the concentration at any stage of the reaction.

The half-life is the time taken for the concentration of a substance to fall to half its initial value. The half-life of a first order reaction is independent of the initial concentration.

Order of Reaction and Rate Constant

You may apply suggested active learning methodologies such as group discussion, brain storming, gaped lectures and problem solving for the delivery of the lessons in this topic. After you give a brief introduction to the topic, have your students perform Activity 4.6. This activity help the students enhance their skills of scientific observation. You are required to guide the students in forming groups, to clarify the questions in the activity where required and to facilitate discussion among group members. Set appropriate time for the activity.

Encourage each student to think independently and share their opinions with their group members. Harmonize the ideas that they develop through discussion, based on the following suggested answers.

If the concentration of a reactant is doubled, the rate of reaction increases twofold. The reactants do not affect the reaction rate in equivalent amounts. The reaction rate does not change in the same ratio every time. In order to understand the nature of a reaction in details chemists need to know the order of a reaction and the rate constant for the reaction.

You may use a problem-based learning methodology and facilitate the discussions by organizing the contents given in the textbook. Make the following points clear to the students.

- the rate of a reaction depends on the concentrations of one or more of the reactants.
- the relationship between the reaction rate and the concentration is given in terms of the rate law or the rate expression.
- the rate law states that the rate is equal to: the rate constant times the concentration of the reactants, raised to simple integers such as zero, one or two.
- the initial rates and initial concentrations of reactions were used to develop the rate law. However, experimental evidence suggests that the rate law is valid for the duration of the reaction.
- if the reaction rate and the concentration are known, then the reaction rate constant can be calculated.
- the unit for the rate constant is in terms of the reciprocal of time, in seconds, minutes, hours etc.

Experimental rate equation

You may apply suggested active learning methodologies such as group discussion, brain storming, gaped lectures and problem solving for the delivery of the lessons in this topic. Let students study Table 4.3 together with the associated text note. You are required to help the students in forming groups so that they may discuss. Appropriate time should also be set for the activity.

Let them think independently and share their opinions with their group members. As you walk around, among the groups and students, observe them and identify the areas of misunderstanding. In your next presentation, resolve their doubts and stabilize the lesson by using Example 4.4.

Arrhenius' equation

Let the students read about this topic, which gives the relationship between E_a, T and k. To motivate their reading, you may set some questions to be answered from the text. Ask the students what problems they faced in understanding their reading assignment. After discussing their questions, use Example 4.5 to stabilize the lesson.

Concentration-Time Equation (Integrated Rate Law)

Introduce this topic by involving students in a discussion of Activity 4.7. This activity is designed to encourage students to express their opinions in the class room. Appropriate time should be set for the activity.

Let them think independently and share their opinions with their group members. Use the following answers as you harmonize the ideas that the students develop through discussion.

Rates of reactions may or may not depend on reactant concentration. See the textbook for detailed explanation. The order of a reaction is obtained from the rate equation.

$$r = k[\mathrm{A}]^x[\mathrm{B}]^y$$

The reaction is x^{th} order in A, y^{th} order in B, and $(x + y)^{th}$ order overall. If the rate of the reaction is independent of the concentration, it is zero order. If the rate is to the first power with respect to concentration of reactant A ($[\mathrm{A}]^1$), then it is first order in A. Similarly, if the rate is to the second power with respect to concentration of reactant A ($[\mathrm{A}]^2$), then it is second order in A. You can also give further explanation.

In your presentation that follows, your approach depends on whether or not your students have studied introductory calculus. Tell the students that concentration is linearly related to time elapsed during the progress of the reaction. Stabilize the knowledge that the students have gained by showing the application of the equation, using Example 4.6. You may give a study assignment on first-order and second-order reactions as homework at the end.

First-order and second-order reactions

You may apply suggested active learning methodologies such as group discussion, brain storming, gaped lectures and problem solving for the delivery of the lessons in this topic. After recalling the main point of the previous lesson, you may begin this topic by evaluating students' ability to understand what they have read from the text. You may raise the following questions:

- How does the concentration-versus-time curve for a first order reaction differ from that of a zero-order reaction?

- How does the concentration-versus-time curve for a second-order reaction differ from that of first order reaction?
- How do the rate equations help us in practical experience? (Relate this question to the examples given.)

Then, you may continue by resolving any doubts the students have. Emphasize the following points:

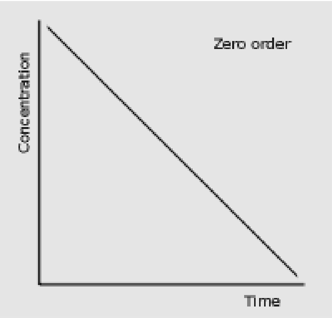
- by using the equations, the reactant concentration at any time during a first-order reaction can be calculated.
- the equations can also be used to find the time taken to reach a specific reactant concentration.

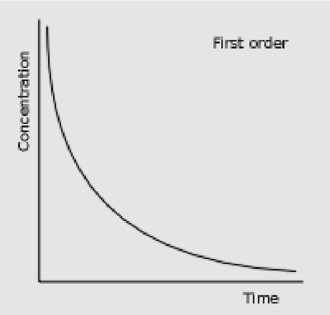
You may refer to the following curves to help the students in comparing the zero-order, first-order and second-order reactions.

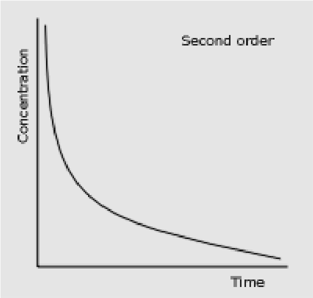
Concentration of a reactant against time

You may apply suggested active learning methodologies such as group discussion and brain storming for the delivery of the lessons in this topic. The plots of the concentration of a reactant against time are given below.

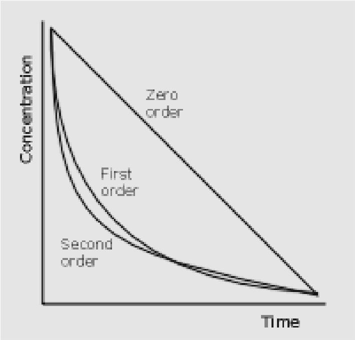
a. zero-order reaction b. first-order reaction c. second -order reaction







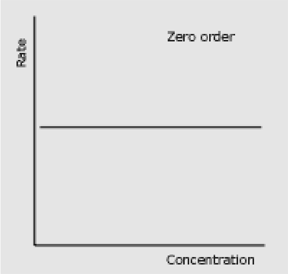
The difference among graphs of the concentration of a reactant against time for zero-order, first-order and second-order reactions can easily be seen below.

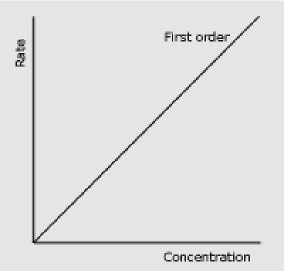


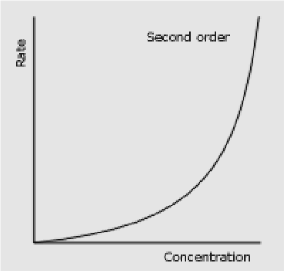
Rate of reaction against concentration of a reactant

You may apply suggested active learning methodologies such as group discussion and brain storming for the delivery of the lessons in this topic. The plots of reaction rate against reactants concentration are given below.

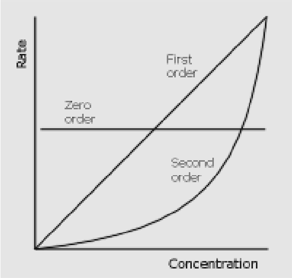
a. zero-order reaction b. first-order reaction c. second -order reaction







The difference among graphs of reaction rate against reactant concentration for zero-order, first-order and second-order reactions can easily be seen below.



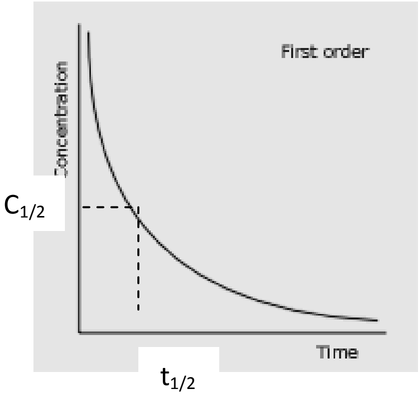
The Half-life of a Reaction

You may apply suggested active learning methodologies such as group discussion and brain storming for the delivery of the lessons in this topic. After introducing the topic, let the students perform Activity 4.8, which gives them practice in plotting rate curves. Encourage the students to read the text note to help them answer the questions. You are required to help the students in forming groups, to clarify the questions in the activity where required and to facilitate their group discussions. Appropriate time should also be set for the activity.

Let them think independently and share their opinions with their group members. Finally, encourage group members to report their common understanding.

After observing the students at work, correct any wrong concepts, using the suggested answer below. Harmonize the ideas developed through discussion, based on the following suggested answers.

A graph that shows the time dependence of reactant concentration for first-order reactions is shown below. On the graph, the time when the concentration of a reactant is decreased by half of its initial amount ($t_{1/2}$) is indicated.



Now, continue the lesson by using the lecture method to explain the steps followed to derive an equation for half-life. In your explanation, help the students to that the concentration of the reactant decreases as the reaction proceeds, because it is being consumed. The time required for the reactant concentration to decrease to one-half of its initial values is called the half-life of a reaction.

After summing up, these concepts have the students read the remaining part of the lesson to prepare for the next period. To motivate give them you may set some questions to answer based on the text.

Application of Half-life

An example involving calculations on the half-life of reactions is given in the student text. Have the students may study the detailed steps taken in the calculations. Your participatory lecture should be aimed at clarifying any point of ambiguity.

Motivate the students to determine the half-life of the first and second order reactions by giving other examples. Use reference books, if available.

Additional Notes

The rate law for a reaction such as:

$$A + B \rightarrow products$$

is represented by

$$Rate = k[A]^x[B]^y$$

where k is the rate constant. The exponents, x and y, are called the order of the reaction with respect to A and B, respectively. The overall order is (x + y). Emphasize the fact that the actual values of x and y can only be obtained by experiment. They have no connection to the coefficient of the balanced chemical equation. If they happen to be the same, then it is mere chance.

For a reaction,

$$A \rightarrow B$$

if doubling the concentration of A doubles the rate, the exponent is 1.

$$\text{Rate} = k[A]^1$$

if doubling the concentration of A quadruples the rate, the exponent is 2.

$$\text{Rate} = k[A]^2$$

where k is the proportionality constant called rate constant. Its values are the rate of a reaction when all the reactant concentrations are 1 M. The rate constant is the one variable that gives the information about the nature of the reactants. Give the students mixed exercises so that they are able to see with values zero and fractions.

According to Arrhenius

$$k = A\ \mathrm{e}^{-Ea/RT}$$

where A is related to the shape and overall nature of the reactant, E_a is the activation energy, T is the temperature in Kelvin, R is the universal gas constant and e is base of natural logarithm. Guide the students helping them to see how big k becomes when A and T are big and E_a low, and vice versa. Limit the discussion to qualitative explanations.

In differential form, the first-order equation is as follows:

$$\frac{-\mathrm{d}[\mathrm{A}]}{\mathrm{d}t} = k[\mathrm{A}]$$

Rearranging, we get $\frac{\mathrm{d}[\mathrm{A}]}{[\mathrm{A}]} = -k\mathrm{d}t$

Integrating between t=0 and t=t gives $\int_{[\mathrm{A}]_o}^{[\mathrm{A}]} \frac{\mathrm{d}[\mathrm{A}]}{[\mathrm{A}]} = -k\int_0^t \mathrm{d}t$

$$\ln\frac{[\mathrm{A}]_t}{[\mathrm{A}]_o} = -kt \text{ or } \ln[\mathrm{A}]_t - \ln[\mathrm{A}]_o = -kt$$

Assessment

Assess students' active involvement in each of the activities. In addition to assessments made by using set questions at the end of each lesson, you may use Exercise 4.7 to assess the achievement of the objective of the lessons in this section.

Your student performance record list is helpful in all your assessments. Appreciate and give extra activities to students working above the minimum requirement level. Assist those working below the minimum requirement level.

The students' performance should be assessed based on whether they are able to:

- express a reaction in terms of the reactant concentration versus time;
- rewrite an algebraic expression for concentrations and time in terms of the equation for a straight-line graph;
- draw graphs, and establish the reaction rate constant from the slope of the graph;
- calculate the time required for a given concentration change;

Evaluate students' ability to understand what they read from the text. Set your own questions to assess the students. You may give them Exercise 4.8 as homework. The students' performance should be assessed based on whether they are able to:

- define the half-life for a first order reaction;
- calculate the time taken for a particular concentration of reactant to be attained; and
- calculate the reactant concentration after a certain time has elapsed.

Additional Questions

*1. Compare the rate in the same reaction, hydrolysis of sucrose ($E_a = 108 \times 10^3$ KJ /at 17 °C) with the rate at 27 $^{\circ}$C.

2. The rate of decomposition of azomethane is studied by monitoring the partial pressure of the reactants as a function of time.

$CH_3\text{-}N{=}N\text{-}CH_3(g) \rightarrow N_2(g) + C_2H_6(g)$

The data obtained at 300°C are shown in the following table.

Time (s)	**Partial pressure of $C_2N_2H_6$ (mm Hg)**
0	284
100	220
150	193
200	170
250	150
300	132

Are these values consistent with first-order kinetics? If so, determine the rate constant.

Answers to Additional Questions

*1. $T_1 = 27 + 273 = 300$ K

$T_2 = 17 + 273 = 290$ K

R = 8.31 J/k.mol

$E_a = 108 \times 10^3 = 1.08 \times 10^5$ kJ /mol

A = not given

We use the same equation as in Exercise 4.7.

$\ln k_1/k_2 = - E_a / RT_1 + Ea / RT_2$

Substituting, we get:

$\ln k_1/k_2 = - 1.08 \times 10^5 / 8.314 \times 300 + 1.08 \times 10^5 / 8.314 \times 290$

$= - 43.300 + 44.794 = 1.493$

$k_1/k_2 = 4.45$

The reaction at 27 °C is more than 4 times faster than at 17 °C.

2. By substituting partial pressure for the concentration,

$\ln[\mathrm{A}]_\mathrm{t} = -kt + \ln[\mathrm{A}]_\mathrm{o}$ can be written as

$\ln P_\mathrm{t} = -kt + \ln P_\mathrm{o}$

First we construct the following table of t vs ln P_t

t (s)	0	100	150	200	250	300
ln P_t	5.649	5.394	5.263	5.136	5.011	4.883

This gives a straight line, so the reaction is indeed first order.

Slope =– 2.55×10^{-3} s^{-1} and $k = 2.55 \times 10^{-3}$ s^{-1}

Answers to Exercise 4.8

1. a. $t_{\frac{1}{2}} = 0.693/k$

$\Rightarrow k = 0.693/t_{\frac{1}{2}} \Rightarrow k = 0.693/30 \text{ min} \Rightarrow k = 0.0231 \text{ min}^{-1}$

b. $$2.303 \log \frac{[A]_0}{[A]} = kt$$

Note: 35 % complete means 65 % remain in the original form.

$$\frac{[A]}{[A]_0}\times 100 = 65,\quad \frac{[A]}{[A]_0} = 0.65 \Rightarrow \frac{[A]_0}{[A]} = \frac{1}{0.65} = 1.54$$

Subsitituting $\frac{[A]_0}{[A]}$ = 1.54 in the above equation,

$$2.303 \log \frac{[A]_0}{[A]} = 2.303 \log (1.54) = kt$$

$$2.303 \log (1.54) = (0.0231 \text{min}^{-1})\, t$$

$$t = 18.70 \text{ min}$$

Note that 35 % is less than half the initial concentration; hence 18.70 min is less than half-life time.

2. a. Substitute the data given in the problem statement into the equation, and then solve for Ea.

$$\log \frac{3.5\times10^{-3}}{2.7\times10^{-4}} = \frac{Ea}{2.303\times8.31 \text{ J mol}^{-1} \text{ K}^{-1})}\left(\frac{1}{600\text{K}} - \frac{1}{650\text{K}}\right)$$

$$1.11 = \frac{Ea}{2.303\times8.31 \text{ J mol}^{-1}} \times (1.28\times10^{-4})$$

Hence,

$$E_a = \frac{1.11\times2.303\times8.31 \text{ J mol}^{-1}}{1.28\times10^{-4}} = 1.66\times10^{5} \text{ J mol}^{-1}$$

Use the same equation, but substitute for K_1, T_1, T_2, and Ea obtained in (a), and solve for k_2. Substitute:

Ea = 1.66 × 105 J/mol and

$k_1 = 2.7 \times 10^{-4}$ L/(mol.s) at $T_1 = 600$ K

k_2 = Unknown as yet at $T_2 = 700$ K

We get

$$\log \frac{k_2}{2.7\times10^{-4}\, L(mol.s)} = \frac{1.66\times10^{5}\, J/mol}{2.303\times8.31 J/(mol.K)} \times \left(\frac{1}{600\text{K}} - \frac{1}{700\text{K}}\right) = 2.07$$

Taking antilogarithms, we obtain

$$\frac{k_2}{2.7\times10^{-4}\, L(mol.s)} = 10^{2.07} = 1.2\times10^{2}$$

Hence $k_2 = (1.2\times10^{2})(2.7\times10^{-4})$ L/(mol.s) $= 3.2\times10^{-2}$ L/(mol.s)

3. The solution to this problem requires solving this equation for Ea

$$E_a = \left(\log\frac{k_2}{k_1}\right)(2.303R)\left(\frac{T_1T_2}{T_2 - T_1}\right)$$

and then just substituting the known values of k_1 and k_2, and T_1 and T_2.

$$E_a = \left[\log\left(\frac{9.30\times10^{-3}\, s^{-1}}{5.54\times10^{-5}\, s^{-1}}\right)\right](2.303)$$

$$\left(8.314\frac{J}{k.mol}\right)\ \frac{(298\,\mathrm{K})(340\,\mathrm{K})}{(340\,\mathrm{K}-298\,\mathrm{K})}$$

$$=(2.23)\left(19.15\frac{J}{k\cancel{.m}ol}\right)(2412.4\cancel{k)}=1.03\times10^{5}\,{}^{J}\!/\!_{mol}$$

4.4 REACTION MECHANISM

Periods Allotted:3 periods

Competencies

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

- *explain reaction mechanism;*
- *explain the molecularity of a reaction;*
- *explain what is meant by rate determining step;*
- *give specific examples to illustrate rate determining step;*
- *explain the relationship between the reaction path way and the rate law;*
- *use rate equation to suggest possible reaction mechanism for a reaction.*

Forward Planning

Take ample time to read in the textbook. Study the concept of the molecularity of an elementary reaction and the reaction mechanism from the textbook. Be prepared to answer questions and solve problems related this subject matter.

Tell the students that, at best, the reaction mechanism which is postulated is acceptable, but that it is not provable.

This is the final topic of the unit. Therefore, let students plan their time for reading the summary and doing the review exercises.

Teaching Aids

No experiment is suggested for this section. Chemical equations written on flipcharts for demonstrating the various lists of elementary reactions are helpful.

Subject Matter Presentation

Suggested active learning methodologies such as brain storming, enquiry, and group discussion may be used for the delivery of the lessons in this topic. This section deals with *reaction mechanism*, which includes the molecularity of elementary reactions (such as unimolecular, bimolecular, termolecular) and the rate-determining step. A reaction mechanism is the step-by-step sequence of elementary reactions by which overall chemical change occurs. It consists of elementary reactions involving individual molecules. An elementary reaction can be classified according to the number of molecules taking part in that particular step (molecularity).

A proposed mechanism must agree with the experimental rate law. The rate law implied by the mechanism is found by identifying the rate-determining step and by finding an expression for the concentrations of any reaction intermediates.

Molecularity of an Elementary Reaction

You may apply suggested active learning methodologies such as group discussion and brain storming for the delivery of the lessons in this topic. Use introductory note in the student textbook to introduce reaction mechanism. Have the students to read this lesson in the classroom and ask some of them to present their understanding to the class.

As you proceed to molecularity of elementary reactions have your students perform *Activity 4.9.* This activity leads to the discussion that follows. Students may read the text note to help them answer some of the questions.

Help the students to come to the following answers. Also, use the answers as you harmonize the ideas the students develope through discussion.

The number of molecular species involved in:

$A \rightarrow products$ is one (unimilecular),

$A + A \rightarrow products$ is two (bimolecular),

$A + B \rightarrow products$ is two (bimolecular),

$\mathrm{A + A + B \rightarrow products}$ is three (termolecular), and

$A + B + C \rightarrow products$ is three (termolecular).

At the beginning of the discussion, be sure that the students understand that one of the objectives of studying reaction rates of chemical reactions is to get an insight into the sequences of chemical steps (reaction mechanism) that occur to produce the overall net reaction. The individual reactions that make up a mechanism are called *elementary steps*. An elementary step is one of the individual steps in a reaction mechanism. It can be one of the following:

- Unimolecular, which involves the decomposition or rearrangement of a single particle
- Bimolecular, which involves the collision of two particles
- Termolecular, which involves the simultaneous collision of three particles all at one time.

We believe that higher molecularities do not occur; because of the law probability. When we have to choose between alternative mechanisms for a reaction, the ones with low molecularity are preferred, all things being equal.

Obtaining a satisfactory mechanism for a reaction is a very difficult task, even for experienced chemists. Therefore, limit your treatment of this subject to qualitative approaches.

Rate Determining Step

You may apply suggested active learning methodologies such as group discussion and enquiry for the delivery of the lessons in this topic. To start the lesson on the rate-determining step, have the students perform Activity 4.10. The purpose of this activity is to encourage the students to express what they think about the rate-determining step and to explain why we need to know about it.

Encourage all of the students to think independently and to share their opinions with their group. Finally, encourage members of their groups to report their common understandings to the class.

After hearing from some group representatives, harmonize the ideas developed through their discussions based on the following suggested answers.

A rate-determining step is the slowest step in the sequence of steps leading to the formation of products. Chemists want to know the rate-determining step because the slowest step in the mechanism determines the overall rate of the reaction.

You may continue your presentation by answering that the students understand that, when there are multiple steps in a reaction mechanism, the slowest step is called *the rate-determining step* because the final products cannot be formed any faster than the products of the slowest step. Be sure that the students know that the predicted rate law only includes the reactants in the overall equation. Neither intermediates nor catalysts appear in the rate law.

Ask your students to do the review exercises.

Additional Notes

Reaction mechanism - elementary process

A *mechanism* for a reaction is a collection of elementary processes (also called elementary steps or elementary reactions) that explains how the overall reaction proceeds.

A mechanism is a proposal from which you can work out a rate law that agrees with the observed rate laws. The fact that a mechanism explains the experimental results is not a proof that the mechanism is correct. A mechanism is our rationalization of a chemical reaction, and devising a mechanism is an excellent academic exercise.

Elementary processes or steps

An *elementary process* is also called an *elementary step or elementary reaction.* It expresses how molecules or ions actually react with each other. The equation in an elementary step represents the reaction at the molecular level, not the overall reaction. Based on numbers of molecules involved in the elementary step, there are three kinds of elementary steps: unimolecular step (or process), bimolecular process, and trimolecular process.

An elementary step is proposed to give the reaction rate expression. The rate of an elementary step is always written according to the proposed equation. This practice is very different from the derivation of rate laws for an overall reaction. When a molecule or ion decomposes by itself, such an elementary step is called a *unimolecular step* (or *process*). A unimolecular step is always a first-order reaction.

The following examples illustrate this point:

$O_3 = O_2 + O \qquad r = k\,[O_3]$

or, in general,

$A = B + C + D \qquad r = k\,[A]$

$A^* = X + Y \qquad r = k\,[A^*]$

* represents an excited molecule.

A *bimolecular process* involves two reacting molecules or ions. The rates for these steps are second order. The following examples illustrate how to give the rate expression. The simulation illustrates a bimolecular process.

$NO + O_3 \rightarrow NO_2 + O_2 \qquad r = k[NO][O_3]$
$Cl + CH_4 \rightarrow HCl + CH_3 \qquad r = k[Cl][CH_4]$
$Ar + O_3 \rightarrow Ar + O_3^* \qquad r = k[Ar][O_3]$
$A + A \rightarrow B + C \qquad r = k[A]^2$
$A + B \rightarrow X + Y \qquad r = k\,[A]\,[B]$

* represents an excited molecule.

A *trimolecular process* involves the simultaneous collision of three molecules. For example:

$O + O_2 + N_2 \rightarrow O_3 + N_2 \qquad r = k\,[O]\,[O_2][N_2]$

$O + NO + N_2 \rightarrow NO_2 + N_2 \quad r = k\,[O][NO][N_2]$

The N_2 molecules in the above trimolecular elementary steps are involved within energy transfer. They cannot be cancelled. They are written in the equation to give an expression for the *rate*s. In general, trimolecular steps may be

$A + A + A \rightarrow \text{products} \qquad r = k\,[A]^3$

$A + A + B \rightarrow \text{products} \qquad r = k\,[A]^2[B]$

$A + B + C \rightarrow \text{products} \qquad r = k\,[A][B][C]$

The simultaneous collision of three molecules is rare, but occasionally reactions take place in this way.

Elementary processes, are written to show how a chemical reaction progresses leading to an overall reaction. Such a collection is called a *reaction mechanism.* In a mechanism, elementary steps proceed at various speeds. *The slowest step is the rate-determining step. The order for that elementary process is the order of a reaction,* but the concentrations of reactants in that step must be expressed in terms of the concentrations of the reactants.

Deriving rate laws from reaction mechanisms

The following example illustrates how elementary steps are used to represent a reaction mechanism. In particular, a slow step in a mechanism determines the rate of a reaction.

Problem 1*:* If the reaction

$2\ NO_2 + F_2 \rightarrow 2\ NO_2F$ follows the mechanism,

i. $NO_2 + F_2 \rightarrow NO_2F + F$ (slow)

ii. $NO_2 + F \rightarrow NO_2F$ (fast)

Work out the rate law.

Solution: Since step i is the rate-determining step, the rate law is

$$-\frac{1}{2}\frac{d[NO_2]}{dt} = k[NO_2][F_2]$$

The addition of (i) and (ii) gives the overall reaction.

Discussion:

This example illustrates that the overall reaction equation has nothing to do with the order of the reaction. The elementary process in the rate-determining step determines the order.

Other possible elementary steps in this reaction are:

$F + F \rightarrow F_2$

$F + F_2 \rightarrow F_2 + F$

$NO_2F + F \rightarrow F + NO_2F$

but they do not lead to the formation of products.

To propose a mechanism requires enough knowledge of chemistry to give plausible elementary processes. A Freshman in chemistry will not be asked to propose mechanisms, but you will be asked to give the rate laws from a given mechanism.

In summary, the number of particles involved in an elementary step is called the molecularity, and, in general, we consider only the molecularity of 1, 2, and 3.

Types of elementary steps are summarized below. In the table, A, B, and C represent reactants, intermediates, or products in the elementary process.

Molecularity	Elementary step	Rate law
1	$A \rightarrow P$	$r = k\,[A]$
2	$A + A \rightarrow P$ $A + B \rightarrow P$	$r = k[A]^2$ $r = k\,[A]\,[B]$
3	$A + A + A \rightarrow P$ $A + 2B \rightarrow P$ $A + B + C \rightarrow P$	$r = k[A]^3$ $r = k[A][B]^2$ $r = k\,[A][B][C]$

where P is products.

Assessment

Devise your own questions to assess the students' overall performance. The students should be able to: define reaction mechanisms as the sequence of bond making and bond- breaking steps that occur during the conversion of reactants to products;

- predict the order of elementary steps by examination of the step equations
- write overall reaction equations from elementary steps, and explain the concept of molecularity, its effect on the rate of elementary steps
- explain why the rate of a reaction depends on the slowest step in the reaction,
- formulate the relationship between the rate and the concentration in terms of the collision theory

You can also use Exercises 4.9 and 4.10 to assess the students' mastery of the section's concepts.

Praise students working above the minimum requirement level and recognize their achievements. Encourage them to continue working hard and not to become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class. Give them extra attention in class and additional lesson time during breaks or at the end of the day.

Additional Questions

*1. The rate law for the reaction

$A + B \rightarrow C + D$

is first-order in [A] and second-order in [B]. If [A] is halved and [B] is doubled, the rate of the reaction will

A. remain the same B. be increased by a factor of 2

C. be increased by a factor of 4 D. be increased by a factor of 8

*2. Bimolecular reactions (e.g., $A + B \rightarrow C$) generally occur at much slower rates than one would predict from the rate of collision between molecules, as calculated from the gas kinetic theory. The discrepancy can be explained in terms of

A. intermolecular repulsions B. non-spherical molecular shapes

C. erroneous estimates of molecular size D. activation energies of reaction

3. If a reaction proceeding by the mechanism

$$A + B \rightarrow C + D$$

occurs at a rate *x*, and if the concentrations of A and B are both doubled, what will be the new rate of reaction?

A. x B. 2x C. 4x D. 8x E. 16x

4. The hypochlorite ion oxidizes the iodide ion in aqueous solution, as represented by the equation

$$OCl^- + I^- \rightarrow OI^- + Cl^-$$

The rate of formation of OI- is given by the rate law

$$r = k\frac{[I^-][OCl^-]}{[OH^-]}$$

The overall reaction order and the order with respect to OH- are

A. 0 and 1 B. 0 and –1 C. 1 and –1 D. 2 and 1

Answers to Additional Questions

*1. B *2. D 3. C 4. C

Answers to Exercise 4.9

1. a. Elementary steps are individual reactions that make up a mechanism.
   b. Unimolecular - involves decomposition or rearrangement of a single particle.
   c. Bimolecular – involves collision of two particles.
   e. Termolecular – involves collision of three particles.
2. a. Unimolecular b. Bimolecular c. Termolecular

Answers to Exercise 4.10

1. The molecularity of a reaction is the number of reactant molecules taking part in a single step of the reaction.
2. The rate-determining step for a reaction that involves more than one step is determined by identifying the slowest step in the reaction mechanism.
3. $NO_2 + NO_2 \rightarrow NO + NO_3$ (slow step) is the rate-determining step. It is bimolecular.

Answers to Review Exercise on Unit 4

Part I: Multiple Choices

1.	B	2.	C	3.	E	4.	C	5.	D
6.	A and B	7.	D	8.	B	9.	C	10.	C
11.	B	12.	C	13.	A	14.	D	15.	A
16.	A	17.	D	18.	A	19.	B	20.	A
21.	B	22.	A	23.	C				

Part II: Answer the following questions briefly

24. a. 2 b. 1 c. 3

25. The order of a reaction is the power to which the reactants or products concentration is raised in the rate law expression. Molecularity of a reaction is the number of species that must collide to produce the reaction.

26. An exothermic reaction gives off (releases) heat energy, whereas an endothermic reaction absorbs heat energy.

27. Pseudo-unimolecular reactions

Consider the following acid-catalysed reactions:

a. Hydrolysis of ethyl acetate

$$CH_3COOC_2H_5 + H_2O + H^+ \rightarrow CH_3COOH + C_2H_5OH$$

b. Inversion of cane sugar

$$C_{12}H_{22}O_{11} + H_2O + H^+ \rightarrow C_6H_{12}O_6 + C_6H_{12}O$$

Both of the above reactions are bimolecular but are found to be of first order. This is because it is experimentally, observed that for the first reaction, the rate of reaction is directly proportional to $[CH_3COOC_2H_5]$ only. For the second reaction, the rate of reaction is directly proportional to $[C_{12}H_{22}O_{11}]$ only. This behavior is due to from the fact that water is present in such large excess that its concentration remains almost constant during the reaction.

Such reactions, which are not truly of the first order, but under certain conditions become reactions of the first order, are known as **pseudo-unimolecular** reactions.

28. Clock reactions are reactions in which the reaction rates can be influenced by factors such as temperature, concentration, catalysts and surface area. When one reactant is consumed or another chemical is produced, an abrupt change in visible color occurs. Since the factors influencing the reaction rate can be changed, the time for the reaction to occur can also be changed thus the idea of a 'clock' reaction.

29. The concentration of A during a reaction of the type

$$A + 2B \rightarrow \text{products}$$

Rate of reaction between each time interval is calculated as follows:

a. $$r = -\frac{\Delta[A]}{\Delta t} = -\frac{(0.229 - 0.317)}{(5 - 1)} = 0.022 \text{mol dm}^{-3} \text{ min}^{-1}$$

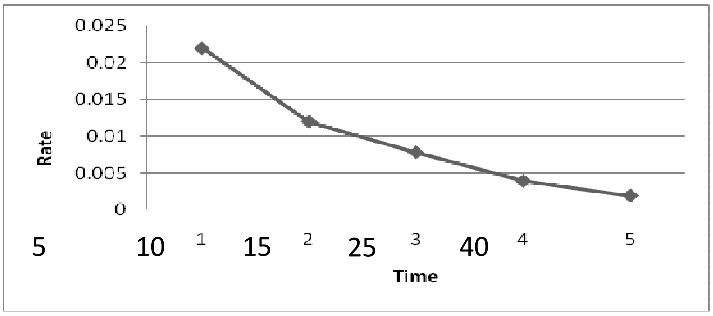
b. $$r = -\frac{\Delta[A]}{\Delta t} = -\frac{(0.169 - 0.229)}{(10 - 5)} = 0.012 \text{ mol dm}^{-3} \text{ min}^{-1}$$

c. $$r = -\frac{\Delta[A]}{\Delta t} = -\frac{(0.130 - 0.169)}{(15 - 10)} = 0.0078 \text{ mol dm}^{-3} \text{ min}^{-1}$$

d. $$r = -\frac{\Delta[A]}{\Delta t} = -\frac{(0.091 - 0.130)}{(25 - 15)} = 0.0039 \text{ mol dm}^{-3} \text{ min}^{-1}$$

e. $$r = -\frac{\Delta[A]}{\Delta t} = -\frac{(0.062 - 0.091)}{(40 - 25)} = 0.0019 \text{ mol dm}^{-3} \text{ min}^{-1}$$

Rate of reaction changes with time, according to the figure below.



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3. http://www.chem.arizona.edu/~salzmanr/480a/480ants/chemkine.html.
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5. http://www.chemlin.de/chemistry/chemical_kinetics.htm.
6. http://www.computationalscience.org/ccce/Lesson10/10b_lab_gides_madonna.pdf
7. http://videolectures.net/mit3091f04_sadoway_lec22.
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9. http://www.chem.purdue.edu/gchelp/116exams/kinetics2.html.
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UNIT 5

CHEMICAL EQUILIBRIUM AND PHASE EQUILIBRIUM

UNIT OVERVIEW Total Periods Allotted: 18 periods

The students were introduced to the concept of chemical equilibrium in grade 9. In this unit, the students will study the concept of chemical equilibrium and phase equilibrium. Therefore, this unit is structured into two sections as follows:

The first section (5.1) deals with the basic concepts of *chemical equilibrium*. In a chemical process, *chemical equilibrium* is the state in which the chemical activities or concentrations of the reactants and products have no net change over time. Usually, this results when the forward chemical process proceeds at the same rate as the reverse reaction.

The reaction rates of the forward and reverse reactions are generally not zero but, being equal, there are no net changes in any of the reactant or product concentrations. This process is called dynamic equilibrium.

In this section, reversible and irreversible reactions, the attainment of equilibrium, equilibrium constant expressions, application equilibrium constant, and Le Chatleries principle are discussed. A chemical equilibrium is a dynamic equilibrium in which the rate of the forward reaction for the formation of the products equals the rate of the reverse reaction for the re-formation of reactants.

The second section (5.2) deals with the concept of phase equilibrium. In the physical sciences, a *phase* is a region of space (a thermodynamic system), throughout which all physical properties of a material are essentially uniform. The term **phase** is sometimes used as a synonym for state of matter. Also, the term *phase* is sometimes used to refer to a set of equilibrium states demarcated in terms of state variables such as pressure and temperature, or by a phase boundary on a phase diagram.

Because phase boundaries relate to changes in the organization of matter such as a change from liquid to solid or a more subtle change from one crystal structure to another this latter usage is similar to the use of "phase" as a synonym for state of matter.

In this section, common terms such as phase components, degree of freedom, phase rule and phase diagram are discussed. A phase is a region of material that is chemically uniform, physically distinct, and (often) mechanically separable. *Gibbs' phase rule* was proposed by Josiah Willard Gibbs in the 1870s as the equality, $F = C - P + 2$, where P is the number of phases in thermodynamic equilibrium with each other, and C is the number of components.

In both sections, many worked examples and exercises are included to provide the students with opportunities to practice. Please advise your students to take the time required to study.

This unit requires a minimum of 18 periods to complete it successfully. We advise you to do the best you can while teaching this unit.

Unit Outcomes

After completing this unit, students will be able to:

- *understand how equilibrium is established;*
- *explain characteristics of dynamic equilibrium;*
- *state the law of mass action and write an expression for the equilibrium constants, K_c and K_p, from a given chemical reaction;*
- *apply the law of mass action to calculate K_c, K_p, concentration and pressure of substances in equilibrium;*
- *understand how the reaction quotient is used to indicate the position of equilibrium;*
- *state Le Chatliers' principle and use it to predict and explain the effects of changes in temperature, pressure, concentration and presence of catalyst on a reaction;*
- *perform an activity to demonstrate the effects of changes in concentration on the position of equilibrium and to determine K_c and K_p values;*
- *explain how equilibrium principles can be applied to optimize the production of industrial chemicals (e.g., production of ammonia and sulphuric acid); and*
- *perform scientific enquiry skills required by this unit: observing, predicting, comparing and contrasting, communicating, asking questions and making generalizations.*

Main Contents

5.1 Chemical equilibrium

5.2 Phase equilibrium

Introduction

This unit is a resource for preparatory school level. It offers a more comprehensive, organized, and measured approach than is found in most standard textbooks. The laws of chemical equilibrium define the direction in which a chemical reaction proceeds. They also define the quantities of reactants and products that will remain after the reaction comes to an end. An understanding of chemical equilibrium and how it can be manipulated is essential for anyone involved in chemistry and its applications.

To begin this unit, use the start-up activity. The purpose of this activity is to draw the attention of the students to concepts which will be discussed in this unit.

You are required to help the students in forming groups, to clarify the questions in the activity where required and to facilitate discussion among group members. Appropriate time should also be set for the activity. After the discussions, encourage group representatives to report the common understanding of their members.

Harmonize the ideas that the students developed through discussion, based on the following suggested answers.

10 mL of 0.1 mol L^{-1} HCl solution is mixed with 10 mL of 0.1 mol L^{-1} NaOH solution. The balanced equation for the reaction between HCl and NaOH is

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

So, the concentration of NaCl formed is 0.1 mol L^{-1}. No HCl and NaOH remains unreacted. Yes, the reaction completes.

According to the equation

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

when 0.1 mol of nitrogen gas and 0.3 mol of hydrogen gas react in a closed two-liter flask, one might expect 0.2 mol of NH_3 to be formed if the reactants are completely consumed to give 100 % yield. But, because of the reversibility of the reaction, less than 0.2 mole of ammonia is actually formed.

5.1 CHEMICAL EQUILIBRIUM

Periods Allotted: 14 periods

Competencies

After studying this section, the students will be able to:

- *explain reversible and irreversible reactions;*
- *define dynamic chemical equilibrium;*

- *state the necessary conditions for attainment of equilibrium;*
- *describe the microscopic event that occurs when a chemical system is in equilibrium;*
- *explain characteristics of chemical equilibrium;*
- *state the law of mass action;*
- *define equilibrium constant;*
- *write the equilibrium constant expression for chemical reactions that involve concentration and partial pressure;*
- *calculate values for equilibrium constants involving concentration and partial pressure;*
- *state the relationship of* K_{eq} *to relative amounts of product and reactants in a given reaction;*
- *show the relationship between* K_c *and* K_p*;*
- *distinguish between homogeneous and heterogeneous reactions;*
- *define reaction quotient;*
- *use the equilibrium quotient to predict the direction of the reaction and the position of equilibrium;*
- *calculate equilibrium concentrations given initial concentrations;*
- *determine whether the reactants or products are favored in ahemical reaction given the equilibrium constants;*
- *list factors that affect equilibrium constants;*
- *state Le-Chatlier's principle;*
- *use Le-Chatlier's principle to explain the effect of changes in temperature, pressure, concentration and presence of catalyst on a reaction;*
- *state the effect of changes in concentration, pressure/volume and temperature on a* K_{eq}*;*
- *perform an activity to demonstrate the effect of changes in concentration and temperature on the position of equilibrium;*
- *perform an activity to determine Kc for esterification organic acids;*
- *define optimum condition;*
- *explain how Le-Chatlier's principle is applied in the Haber process of manufacturing ammonia and contact process of manufacturing sulphuric acid.*

Forward Planning

Read and understand the notes included in the textbook. In addition to this, read about chemical reaction in the chemistry grade 9 textbook in order to understand the students' background knowledge related to this topic. Schedule the time required for the activities suggested in this subtopic. Be sure that all required materials are ready for the activities suggested in this section.

You need to read and understand the notes included in the students' text. You may also enrich your understanding by reading other reference books.

You may need to prepare the figures in this section on flip charts.

Study the calculations performed in the examples given.

Be sure that the equipment and chemicals required for Experiment 5.1 are available, and schedule the laboratory session.

Perform the experiment by yourself before the actual lesson session. This helps you to become familiar with the procedure and to find out the length of time required for the reaction to attain equilibrium position. You may assign students the task of reading the content in the student text ahead of time.

Teaching Aids

Refer to the students' textbook for the apparatus and chemicals required for performing experiment 5.1–5.3. Diagrams prepared on flip charts for the demonstration of chemical equilibrium are also helpful to the students.

Subject Matter Presentation

Active learning methodologies such as group discussion, brain storming, problem solving and spider diagram are generally suggested for the delivery of the lessons in this topic. This section includes presentation of the concept of chemical equilibrium with major emphasis on reversible and irreversible reactions, attainment and characteristics of chemical equilibrium, law of mass action, equilibrium constant, equilibrium concentration, and reaction quotient.

Reversible and Irreversible Reactions

You may apply suggested active learning methodologies such as group discussion, brain storming and problem solving for the delivery of the lessons in this sub-topic. Chemical equilibrium, was introduced in Grade 9. Here the subject is treated in greater depth. To start the unit, let students discuss *Activity 5.1.* This activity is designed to encourage the students to understand the reversible process that takes place when ice is heated and cooled. You are required to help the students in forming groups, clarify the questions in the activity where required and facilitate the discussions among group members. Appropriate time should also be set for the activity.

Let them think independently and share their opinions with their group members.

Finally encourage representatives of various groups to report the common understanding of group members.

The following answers help you to harmonize the ideas developed through discussion.

In a test tube, crushed ice was taken. A thermometer was dipped in it and the temperature was noted. It was heated slowly for a few seconds. Heating was stopped and the test tube was replaced in a beaker containing large amount of crushed ice.

This activity involves phase changes between ice and liquid water. When water is cooled, it freezes and forms ice. When ice is heated it melts to form liquid water.

$$\underset{\text{ice}}{H_2O(s)} \rightleftharpoons \underset{\text{liquid water}}{H_2O(l)}$$

The temperature of crushed ice is 0°C, the temperature at which water freezes to form an ice. When crushed ice is heated slightly, it gradually starts melting. When it is cooled again, it is changed into ice. The final temperature is expected to be 0°C.

Let students describe reversible reactions and give the example of the physical process already discussed. Next, let them perform *Activity 5.2*. The purpose of this activity is to teach the student how to perform the practical activity of categorizing a given solution as acid or a base. You are required to help the students in forming groups, to clarify the procedures followed in the activity where required and to provide them with the equipment and chemicals required for the activity.

Encourage them to think independently and share their ideas. Harmonize the ideas developed through discussion, based on the following suggested answers.

1. Placing of a drop of 0.1 M HCl solution on litmus paper changes it from blue to red.
2. Placing a drop of 0.1 M NaOH solution on litmus paper changes it from red to blue.
3. When 10 mL each of 0.1 M HCl and 0.1 M NaOH solutions are mixed and tested with litmus paper, no color change is observed.
4. When a pinch of sodium chloride is dissolved in 10 mL water, the resulting solution is neutral.

Based on these observations, help the students to arrive at the following generalizations:

- acids change the color of litmus paper from blue to red.
- bases change the color of litmus paper from red to blue.
- Mixing HCl and NaOH solutions gives a neutral solution.
- Dissolving sodium chloride in water produces a neutral solution.

- Comparing the results of 3 and 4, we see that in both cases neutral NaCl solution is formed.

You may summarize the concept developed in your final mini-lecture. Note that students might easily form misconceptions above chemical equilibrium. Therefore, give a thorough explanation of reversible and irreversible reactions, with the help of examples.

Attainment and Characteristics of Chemical Equilibrium

Group discussion, brain storming, and spider diagram are suggested as active learning methodologies for the delivery of the lessons in this sub-topic. Here again, students are given opportunity to discuss the questions in the next activity. *Activity 5.3* helps students understand the equilibrium of acetone molecules between the liquid and vapor phases. You are required to help the students in forming groups, clarify the procedures followed in the activity where required and provide them with the equipment and the chemicals required for the activity.

Encourage them to think independently and share their ideas.

The following answer helps you to harmonize the ideas suggested by the students.

Acetone is a well-known organic solvent in chemical laboratories. It is a volatile solvent. Initially, 10 mL of acetone is put into a test tube, and the mouth of the test tube is closed immediately. When the test tube is clamped and allowed to stand vertically for half an hour, the level of the solvent in the test tube decreases at regular intervals. Since the test tube is closed, there is no escape of acetone vapors to the external environment. However, some acetone molecules leave the liquid state and escape into the space above the liquid, and equilibrium is established between acetone molecules in both states.

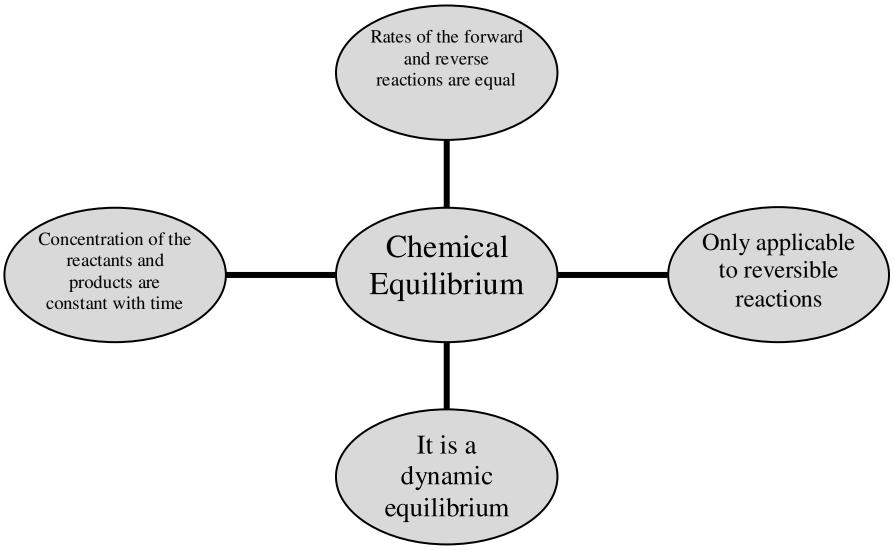
$$CH_3COCH_3(l) \rightleftharpoons CH_3COCH_3(g)$$

After harmonizing the discussions, you may continue to elaborate the state of chemical equilibrium. Making use of figures drawn on flip charts in your discussion is helpful.

The lesson may be started by applying the concept-mapping methodology in order to help students begin to understand interrelationships among concepts. This method is suitable for developing students' cognitive structures. It can also contribute significantly towards meaningful learning in the mind of the student.

Devise a mechanism to help students appreciate that chemical equilibrium applies only to reversible chemical reactions and is a dynamic one (not static). The rates of both the forward and reversible reactions are equal in a chemical reaction at equilibrium. Hence, no shift in position of equilibrium is observed.

After the students make their own attempts at concept mapping, you may use the spider diagram as shown below.



Exercise 5.2 may be given as class work to help you assess your students' achievements.

Law of Mass Action (Equilibrium Law) Expression

You may use suggested active learning methodologies such as group discussion, brain storming and problem solving for the delivery of the lessons in this sub-topic. You may assign your students the task of reading the explanation given in their text on *law of mass action* and *rules for writing the equilibrium constant expression for homogeneous reactions.* Then, you may let the students discuss what they have read.

In your presentation, you are required to harmonize the ideas developed by the students with facts that they should understand in order to meet the objective of the lesson. Your elaboration should be well supported by examples. They must be able to see that the equilibrium constant is a special reaction quotient, which itself is a numerical value of the mass action expression that is determined when the system is at equilibrium.

After assigning Exercise 5.3 to students as homework, you may continue your lesson by presenting the units of K_c and K_p. You may facilitate the discussion by encouraging the students to become involved by asking and answering questions.

Equilibrium Constant for Heterogeneous Reactions

Active learning methodologies including group discussion, brain storming and problem solving are suggested for the delivery of the lessons in this sub-topic Begin this lesson with an introductory statement of the topic. Then, have your students discussed *Activity*

5.4. This activity is designed to encourage the students to identify the phases of reactants and products and to name such type of equilibrium.

During their discussion, let them focus on key terms such as concentration, rate of reaction, homogenous and heterogeneous reactions, constant terms, etc. Let them read their textbook to get helpful ideas for their discussion. Don't expect students to give exact answers to the questions.

The following answers help you to harmonize the ideas developed through discussion.

Consider the reaction

$$Mg(s) + HCl(aq) \rightleftharpoons MgCl_2(aq) + H_2(g)$$

1. Mg is in a solid phase, HCl is in aqueous phase, $MgCl_2$ is in aqueous phase and hydrogen is in a gas phase. Therefore, the reactants and the products are in different phases.
2. The equilibrium that involves reactants and products in different phases is called heterogeneous equilibrium.

The next discussion concerns on rules for writing the expression for the equilibrium constant for heterogeneous reaction. You can facilitate it by letting students read the notes given and point out the different issues that are relevant to such types of reactions. Let the students use the examples in their discussion.

Exercise 5.4 may be given as homework. The lesson on the relationship between K_c and K_P may be handled in a similar way.

Applications of Equilibrium Constant

As you deal with the lesson on *applications of equilibrium constant*, let the students discuss the issue in *Activity 5.5*. The purpose of this activity is to encourage the students to develop the concept of direction of shift of reaction to reach equilibrium. You are required to help the students in forming groups, to clarify the discussion points in the activity, where required, and to facilitate their discussion. Let students think independently and write down their answers in their notebooks. Next, let them share their opinion with their friends in their groups. Some of them may move to other groups to share the work of their group with students in the other group. Finally, let some students from different groups give their answers to the whole class.

Use the following answers the ideas suggested by the students.

Inserting the starting concentrations in the equilibrium constant expression, we write

$$\frac{[HI]_o^2}{[H_2]_o[I_2]_o} = \frac{(1.98)^2}{(0.243)(0.146)} = 111$$

Because the quotient $\frac{[HI]_o^2}{[H_2]_o[I_2]_o}$ is greater than K_c, this system is not at equilibrium. Consequently, some of the HI will react to form more H_2 and I_2 (decreasing the value of the quotient). Thus, the net reaction proceeds from right to left to reach equilibrium.

After delivering a mini-lecture on description given in the textbook accompanying this concept, ask the student to perform *Activity 5.6* individually. This activity is prepared to give students chemical information that can be obtained from an equilibrium constant, K. Encourage them to think independently. Let them form groups and share their ideas. Complement the ideas developed through discussion, based on the following suggested answers.

1. Given the equations

   a $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $K_c = 3.6 \times 10^8$

   the formation of the product is favored, since K_c is extremely large.

   b. $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $K_c = 5 \times 10^{-3}$

   the formation of the product is not favored, since Kc is small.

   c. $CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$ $K_c = 4.45 \times 10^{-24}$

   the formation of the product is favored, since K_c is extremely small.

2. Increasing tendency to proceed towards completion (least extent to greatest extent).

   b. $K_c = 2.1\times10^{-4}$ < (c) $K_c = 4.7\times10^{-4}$ < (a) $K_c = 13.8$

Next, present a brief lesson continues on *reaction quotient and its application.* This may be given to students as a reading assignment. You are required to prepare some key questions to motivate students to read and understand. Don't forget to assess whether the students have performed the task.

Calculating the Equilibrium Concentrations

Active learning methodologies such as group discussion, brain storming and problem solving are suggested for the delivery of the lessons in this sub-topic. This lesson may be given to students to read and discuss in groups. You are required to facilitate the discussion. While walking around the classroom, discussion groups, and identify the student doubts and the ambiguities in their ideas. In your presentation, make the ideas clear in brief. The presentation of this topic needs to be supported with Experiment 5.1. It is designed to help students acquire the practical skills of determining equilibrium constants for the esterfication of acetic acid. Refer to the students' textbook for the detailed procedures as well as for the apparatus and chemicals required for performing this experiment.

This experiment cannot be completed within a period since the procedure requires refluxing time of 1 hour. Therefore, if possible arrange a special time for the experiment after school or during week days. If this is not possible in your school situation, you may first start the refluxing procedure before you invite the students to the laboratory and demonstrate the rest of the procedure in the remaining time.

In order to demonstrate this experiment to the students, use the *predict-observe-explain* methodology**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question, which in turn gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

The amount of acetic acid consumed is equal to the amount of propan-1-ol consumed. It is also equal to the amount of ester formed. The amount of acetic acid that remains at equilibrium will be equal to the amount of NaOH required during titration. During the titration, you may use the following table to record the volume of 0.01 M NaOH used to produce an organized set of data.

Trial No.	***Initial volume of NaOH (mL)***	***Final volume of NaOH (mL)***	***Volume used (mL)***
1			
2			
3			

Esterification of acetic acid is represented by the equation:

$$CH_3COOH\ (aq) + CH_3CH_2CH_2OH\ (aq) \rightleftharpoons H_2O\ (l) + CH_3COOCH_2CH_2CH_3\ (aq)$$

Neutralization during titration:

$$CH_3COOH\ (aq) + NaOH\ (aq) \rightarrow CH_3COONa\ (aq) + H_2O\ (l)$$

Suppose the volume of NaOH used for complete titration of acetic acid to be 10 mL.

$$\text{No. of moles NaOH} = 0.1\ M \times 10/1000\ L = 1.00 \times 10^{-3}\ mol$$

which is also equal to the number of moles of CH_3COOH that remains unreacted.

$$CH_3COOH\ (aq) + CH_3CH_2CH_2OH\ (aq) \rightleftharpoons CH_3COOCH_2CH_2CH_3(aq) + H_2O(l)$$

Initial: 5×10^{-3} mol 5×10^{-3} mol 0 0

At equilibrium: 1.00×10^{-3} mol 1.00×10^{-3} mol 4×10^{-3} mol 4×10^{-3} mol

Assuming the total volume of the solution is 20 mL at the end, the *concentration of the individual species at equilibrium* can be calculated as follows:

$[CH_3COOH] = 1.00 \times 10^{-3}\ \text{mol}/20 \times 10^{-3}\ \text{L} = 0.05\ \text{M}$

$[CH_3CH_2CH_2OH] = 1.00 \times 10^{-3}\ \text{mol}/20 \times 10^{-3}\ \text{L} = 0.05\ \text{M}$

$[CH_3COOCH_2CH_2CH_3] = 4 \times 10^{-3}\ \text{mol}/20 \times 10^{-3}\ \text{L} = 0.2\ \text{M}$

The equilibrium constant for the esterification reaction is calculated as follows.

$$k_c = \frac{[CH_3COOC_3H_7]}{[CH_3COOH][C_3H_7OH]} = 0.2\ \text{M} /(0.05\ \text{M} \times 0.05\ \text{M}) = 80\ \text{M}^{-1}$$

In your experiment, if you get a different amount of NaOH consumed in the titration, you may do the calculation in a similar way. You have to keep on refluxing until the volume of the NaOH needed for complete titration is more or less constant. This may be checked by taking a 1 mL sample of the reaction mixture at some time intervals and titrating it with NaOH. If the amount of NaOH consumed is constant, you will be sure that the reaction has reached the equilibrium position. If the volume of NaOH needed is not constant, it means the concentration of ethanoic acid is still decreasing, and the forward rate is larger than the backward rate. We need to keep on refluxing unitl the forward and backward rates become the same.

The answers to the questions listed under observation and analysis are given below.

1. The purpose of adding a small amount of concentrated H_2SO_4 to the reaction mixture at the beginning of the experiment is to use it as a positive catalyst to speed up the reaction by providing an alternate pathway with lower activation energy. It was used to correct the volume of CH_3COOH.

2. We need to add boiling chips to the reactant mixture before refluxing to prevent the solution from boiling vigorously during heating. Also, propan-1-ol is a volatile solution. It will easily catch fire if it is boiled vigorously.

Changing Equilibrium Conditions - Le Chatelier's Principle

You may apply suggested active learning methodologies such as group discussion, brain storming and peer-teaching for the delivery of the lessons in this sub-topic. *Activity 5.7* is designed to introduce this topic. Appropriate time should be set for the activity. You are required to help your students to understand the questions. Let them think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. The following answer helps you to harmonize the ideas developed through discussion.

The technician should have known that a catalyst cannot change the position of equilibrium. A catalyst lowers the activation energy and so speeds up the rate of both the forward and reverse reactions and does not change the equilibrium constant. Equilibrium is established in the reversible reaction

$$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g) \qquad \Delta H = 114.4\text{ kJ}$$

The changes that could be made to this mixture to increase the amount of $Cl_2(g)$ at equilibrium include:

- increasing temperature, since the forward reaction is endothermic
- removing $Cl_2(g)$ as soon as it is formed
- increasing the concentration of one of the reactants
- increasing pressure

shift the equilibrium position to the right so that the amount of $Cl_2(g)$ at equilibrium increases.

A reaction having a small value of K_c to completion can be derived to completion by altering the factors which affect the equilibrium such as concentration, temperature, pressure, etc.

In your explanation that follows, help your students describe Le Chatelier's principle, as applied to chemical equilibrium. One common point confused here is that after a new equilibrium is established, right after subjecting an equilibrium to stress, the equilibrium constant remains unchanged, but the individual concentration of reactants and products could change.

Effect of Change in Temperature on the Position of Equilibrium

Group discussion, brain storming and peer-teaching are suggested active learning methodologies for the delivery of the lessons in this sub-topic. This lesson may be arranged in a lab session. The purpose of *Activity 5.8* is to promote scientific observation by the students and their ability to give reasons. You are required to help the students in forming groups, clarify the procedures followed in the activity, where required, and provide them the equipment and chemicals required for the activity. Let them think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. The following answers help you to harmonize the ideas developed through discussion.

Some ionic compounds exist as hydrates. They form weak bonds to water molecules. The attachment of these water molecules can affect the electronic structure of the compound and affect its color. An example is cobalt(II) chloride. Without the attached water molecules cobalt(II) chloride is a blue solid. When exposed to humid air, however, the salt forms a hydrate and turns a dark red. The compound is called cobalt(II) chloride hexahydrate, and its formula is $CoCl_2{\cdot}6H_2O$.

This process can be represented by equation:

$$CoCl_2\,(s) + \quad 6H\ O(g) \quad \rightarrow \quad CoCl_2.6H_2O(s)$$

blue from air dark red

By heating the hydrate, water can be driven off:

$$CoCl_2\ .6H\ O(s) \underset{cool}{\overset{heat}{\rightleftarrows}} CoCl\ (s) \quad + \quad 6H\ O(g)$$

dark red blue

When dissolved in water, the cobalt(II) chloride salt decomposes, resulting in the formation of the $Co(H_2O)_6^{2+}$ ion and a deep red solution.

$$CoCl_2.6H\ O(s) \xrightarrow{H_2O(l)} Co(H_2O)_6^{2+}(aq) + 2Cl\ (aq)$$

Dark red crystals red solution

Alternatively, solutions high in chloride concentration can form the dark blue aqueous $CoCl_4^{2-}$ ion.

$$Co(H_2O)_6^{2+}(aq) + \quad 4Cl\ (aq) \rightleftarrows \quad CoCl_4^{2-}(aq) + \ 6H\ O(l)$$

blue

The discussion that follows can be facilitated by involving students in questions and answers activities. You may ask your students about the effect of temperature on equilibrium position in exothermic and endothermic reactions, using the specific examples mentioned in the student text.

Experiment 5.2 is designed to help students understand the effect of changes in temperature on the equilibrium position of starch and iodine. Refer to the students' textbook for the detailed procedures, as well as for the apparatus and chemicals required for performing this experiment.

You may use the predict-observe-explain strategy in order to demonstrate this experiment to the students**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question, which in turn gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

When a few drops of tincture of iodine were added to some starch solution, a blue-black color solution was formed. When the solution was heated to about $80^{o}C$, the blue-black color disappeared.

The answers to the questions listed under observation and analysis are given below.

1. When iodine reacts with starch, starch-iodine complex is formed. This is an example of equilibrium reaction.
2. The forward reaction, which is in the direction of starch-iodine complex formation, is exothermic. The reverse is endothermic as shown below.

$$\text{Iodine} + \text{starch} \underset{\text{endothermic}}{\overset{\text{exothermic}}{\rightleftharpoons}} \text{Starch-Iodine complex}$$

3. As we apply to the system under equilibrium, the equilibrium position shifts in the direction that favors the reactant side. This is because the forward reaction is exothermic. As a result of this, the blue-black color gradually disappears and the solution becomes colorless.

The students may be assigned the task of reading and studying the lesson that follows: "*Effect of Change in Pressure on the Position of Equilibrium*". You are required to provide them with some key questions to be answered after they have studied the lesson. In the next period, you may assess your students' achievements by asking the questions.

Effect of Change in Concentration on Equilibrium Position

Active learning methodologies such as group discussion, brain storming and peer-teaching are suggested for the delivery of the lessons in this sub-topic. This subject may also be treated in a lab session. *Activity 5.9* is helpful for developing the concept of equilibrium. You are required to help the students in forming groups, clarify the questions in the activity, where required, and facilitate their discussion among group members. Appropriate time should also be set for the activity.

Let them think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. Harmonize the ideas developed through discussion, based on the following suggested answers.

Radioactive ^{131}I will find its way into the I_2 molecule, but all of it will not form molecules of $I^{131}I$. Some will remain as $H^{131}I$. First of all, ^{131}I will react just like any other iodine atom. The reaction system is at equilibrium, which is a dynamic equilibrium. Therefore, the forward and reverse reactions are occurring simultaneously and at the same rate. Any $H^{131}I$ will eventually decompose by the reverse reaction and form $I^{131}I$ molecules. Not all of the ^{131}I will end up in the $I^{131}I$ molecules because the forward reaction converts $I^{131}I$ back into $H^{131}I$. Therefore, the ^{131}I will be distributed between the two iodine-containing molecules.

Experiment 5.3 to helps students study the effect of changes in concentration on equilibrium position. Refer to the students' textbook for the detailed procedures as well as for the apparatus and chemicals required for performing this experiment.

You may apply the predict-observe-explain strategy in order to demonstrate this experiment to the students**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question, which in turn gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

The students combine $Fe(NO_3)_3$ and KSCN (potassium thiocyanate potassium is a spectator ion) to produce the red-colored solution of $FeSCN^{2+}$, Iron(III)-thiocyanate.

This is the main reaction that the students consider.

$$\mathrm{Fe^{3+}(aq)} \quad + \quad \mathrm{SCN^{-}(aq)} \rightleftharpoons \mathrm{FeSCN^{2+}\,(aq)}$$

colorless colorless red-brown

$$Kc = \frac{[\mathrm{FeSCN^{2+}}]}{[\mathrm{Fe^{3+}}][\mathrm{SCN^{-}}]}$$

The answers to the questions listed under observation and analysis are given below.

When the concentration of $Fe(NO_3)_2$ increases, by Le Chatelier's principle, the position of equilibrium shifts to the right so as to remove some of the extra $SCN^-(aq)$. More $FeSCN^{2+}(aq)$ produced, and the solution becomes darker red.

To help students sort out the e*ffects of other factors on the position of equilibrium and the optimum conditions for a reaction, you may* give it to them as a reading assignment and let them present their findings to their classmates. Hence, the assignment must be announced ahead of time. Some questions that can be answered from the specified passages are helpful for the preparation of the students.

Chemical Equilibrium and Industry

You may use the case-study method of teaching so that the students may study in groups and present to the class. Hence, give a reading assignment to the students ahead of time. You may prepare your own activity for this lesson. Let the students read the content and discuss among groups. Help the students to see the application of Le Chatelier's principle in the industrial production of important substances such as ammonia and sulfuric acid. Emphasize experimental parameters such as temperature, pressure and concentration that may affect the percentage yield of the products. Facilitate their discussion and harmonize the facts. You may use a flow scheme prepared on a flip chart.

Additional Notes

Definition of Chemical Equilibrium

Chemical equilibrium applies to reactions that can occur in both directions. In a reaction such as:

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$$

The reaction can happen both ways. So, after some of the products are created, the products begin to react to form as the reactants.

At the beginning of the reaction, the rate at which the reactants are changing into the products is higher than the rate the products are changing into the reactants. Therefore, the net change is a higher number of products.

Even though the reactants are constantly forming products and vice-versa, the amount of reactants and products does become steady. When the net change of the products and reactants is zero, the reaction has reached equilibrium.

This equilibrium is a dynamic equilibrium. The definition for a dynamic equilibrium is when the amounts of products and reactants are constant. Note: They are not equal but constant. Also, both reactions are still occurring.

Equilibrium Constant

To determine the amount of each compound that will be present at equilibrium, you must know the equilibrium constant. To determine the equilibrium constant, you must consider the generic equation:

$$aA + bB \rightleftharpoons cC + dD$$

The uppercase letters are the molar concentrations of the reactants and products. The lowercase letters are the coefficients that balance the equation. Use the following equation to determine the equilibrium constant (K_c).

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

For example, determining the equilibrium constant of the following equation can be accomplished by using the K_c equation.

Using the following equation, calculate the equilibrium constant.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

A one-liter vessel contains 1.60 moles NH_3, 0.800 moles N_2, and 1.20 moles of H_2. What is the equilibrium constant?

$$K_c = \frac{(1.60)^2}{(0.80)(1.20)^3} = 29.6 \qquad \text{Answer: 1.85}$$

Le Chatelier's Principle

Le Chatelier's principle states that when a system in chemical equilibrium is disturbed by a change of temperature, pressure, or concentration, the system shifts in equilibrium composition in a way that tends to counteract this change of variable. The three ways that Le Chatelier's principle says you can affect the outcome of the equilibrium are as follows:

i. Changing concentrations by adding or removing products or reactants to the reaction vessel.

ii. Changing partial pressure of gaseous reactants and products.

iii. Changing the temperature.

These actions change each equilibrium differently. Therefore, you must determine what needs to happen for the reaction to get back in equilibrium.

Example involving change of concentration:

In the equation

$$2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$$

If you add more $NO_{(g)}$, the equilibrium shifts to the right, producing more $NO_{2(g)}$. If you add more $O_{2(g)}$, the equilibrium shifts to the right, producing more $NO_{2(g)}$. If you add more $NO_{2(g)}$, the equilibrium shifts to the left, producing more $NO_{(g)}$ and $O_{2(g)}$.

Example 1: pressure change

In the equation,

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

an increase in pressure will cause the reaction to shift in the direction that reduces pressure, which is toward the side with fewer gas molecules. Therefore, an increase in pressure will cause a shift to the right, producing more product. (A decrease in volume is one way of increasing pressure.)

Example 2: temperature change

In the equation,

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_3 + 91.8\text{ kJ}$$

an increase in temperature will cause a shift to the left because the reverse reaction uses the excess heat. An increase in forward reaction would produce even more heat since the forward reaction is exothermic. Therefore, the shift caused by a change in temperature depends upon whether the reaction is exothermic or endothermic.

Assessment

You are required to assess students' active involvement in each of the activities. Your student-performance record list is helpful in all your assessments. Appreciate and give extra activities to students working above the minimum requirement level. You are required to assist those working below the minimum requirement level.

Check that students have developed the calculation skills recommended in the section.

Exercises 5.1 through Exercise 5.7 may help you to assess the achievement of the objectives. Some of them may be assigned as class work and others as homework. You may also modify these questions in order to assess student's level of understanding in a simple way.

Set questions to evaluate whether students are able to:

- explain the characteristics of irreversible and reversible reactions;
- describe the characteristics of a reaction at chemical equilibrium;
- describe the dynamic nature of chemical equilibrium; and
- explain the need for a closed system for a reaction to reach equilibrium.

Additional Questions

On the line at the left, write the letter of the description that best matches each term.

__*1. Equilibrium position
__*2. Law of chemical equilibrium
__*3. Reaction quotient
__*4. Law of mass action
__*5. Equilibrium constant

a. used to determine whether a reaction has reached equilibrium.
b. depends on the initial concentrations of the substances in a reaction.
c. states that every reaction proceeds to an equilibrium state with a specific K_{eq}.
d. expresses the relative concentration of reactants and products at equilibrium in terms of an equilibrium constant.
e. the ratio of product-concentration to reactant-concentration at equilibrium.

*6. a. Write the equilibrium constant expression, K_c, for the reaction below.

$$3\,Fe(s) + 4\,H_2O\,(g) \rightleftharpoons Fe_3O_4\,(s) + 4\,H_2\,(g)$$

b. For the reaction,

$$2\ NOCl\ (g) \rightleftharpoons 2\ NO\ (g) + Cl_2\ (g)\ K_c \text{ is } 2.4 \times 10^{-7}.$$

Caculate the equilibrium constant, K_c, for the reaction

$$1/3\ Cl_2\ (g) + 2/3\ NO(g) \rightleftharpoons 2/3\ NOCl\ (g)$$

c. Given the following equilibrium equations and their corresponding equilibrium constants:

$$2CO_2\ (g) + H_2O\ (g) \rightleftharpoons 2O_2\ (g) + CH_2CO\ (g) \qquad Kc = 6.1 \times 10^{8}$$

$$CH_4\ (g) + 2O_2\ (g) \rightleftharpoons CO_2\ (g) + 2H_2O\ (g) \qquad Kc = 1.2 \times 10^{14}$$

Find Kc for the reaction

$$CH_4\ (g) + CO_2\ (g) \rightleftharpoons CH_2CO\ (g) + H_2O\ (g)$$

7. A mixture of 9.22 moles of A, 10.11 moles of B, and 27.83 moles of C is placed in a one-liter container at a certain temperature. The reaction is allowed to reach equilibrium. At equilibrium, the number of moles of B is 18.32. Calculate the equilibrium constant for the reaction:

$$A\ (g) + 2\ B\ (g) \rightleftharpoons 3\ C\ (g$$

8. a. At a certain temperature, Kc is 4.13×10^{-2} for the equilibrium:

$$2\ IBr\ (g) \rightleftharpoons I_2\ (g) + Br_2\ (g)$$

Assume that equilibrium is established at the above temperature by adding only IBr (g) to the reaction flask. What are the concentrations of I_2 (g) and Br_2 (g) in equilibrium with 0.0124 moles/liter of IBr(g) ?

b. What was the initial concentration of IBr before equilibrium was es tablished?

9. 0.924 mole of A(g) is placed in a 1.00 liter container at 700 °C, where it is 38.8 % dissociated when equilibrium is established.

$$3\ A(g) \rightleftharpoons 5B(g) + 2\ C(g)$$

What is the value of the equilibrium constant, K_c, at the same temperature?

Answers to Additional Questions

*1. e *2. c *3. b *4. d *5. a

*6. a. $[H_2]^4/[H_2O]^4$

b. 1.6×102

c. 7.3×1022

7. The equilibrium constant for the reaction:

   $A (g) + 2 B (g) \rightleftharpoons 3 C (g)$ is 0.832

8. a. The concentrations of I_2 (g) and Br_2 (g) in equilibrium with 0.0124 moles/liter of IBr(g) is 2.52×10^{-3} M

   b. The initial concentration of IBr before equilibrium was established is 0.0174.

9. The value of the equilibrium constant, Kc, at the same temperature is 0.0241.

Answers to Exercise 5.1

1. A reversible reaction is a reaction in which simultaneously, the reactants form products and the products form to give back the reactants. Chemical reactions which proceed in only one direction are known as irreversible reactions.

2. The formation of ammonia is an example of a reversible reaction,

   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   And the burning of magnesium is an irreversible reaction.

   $Mg(s) + O_2(g) \rightarrow MgO(s)$

   Students may also give other examples.

3. The symbol or sign of a single arrow (→) is used to indicate that the reaction is irreversible and a double arrow (⇌) is used to indicate that the reaction is reversible.

Answers to Exercise 5.2

1. a. When a mixture of SO_2 and O_2, for instance, is introduced into a reaction vessel at a temperature of 700 K, a reaction that produces SO_3 occurs:

   $$2SO_2 + O_2 \rightarrow 2SO_3$$

   When a specific concentration of SO_3 is reached, no additional SO_3 is formed, even though some SO_2 and O_2 remain. From that time on, the concentrations of SO_2, O_2, and SO_3 stay constant, and we say the system has reached chemical equilibrium. The constant concentrations are the result of a reversible chemical reaction. In the reverse reaction, some SO_3 decomposes back into SO_2 and O_2.

   $$2SO_3 \rightarrow 2SO_2 + O_2$$

   When the rates of forward and reverse reactions are the same, no net chemical change occurs, and a state of chemical equilibrium exists.

b. The equilibrium state is referred to as dynamic because of the continual conversions of reactants into products, and products into reactants, at the molecular level. A reversible reaction is represented by the opposing arrows in the chemical equation:

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

c. This equation shows three compounds, SO_2, O_2 and SO_3 in equilibrium:

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

At 700 K, all three species exist together in the reaction vessel. Equilibrium can be reached by the reverse reaction as well. That is, if only SO_3 is added to the reaction vessel at 700 K, it will form an equilibrium amount of SO_2 and O_2.

2. In the given figure (a) rate of forward reaction (b) rate of reverse reaction, and (c) equilibrium point. The rate of the forward reaction decreases and the rate of the reverse reaction increases.

3. In the given figure (a) Concentration of product (b) concentration of reactants, and (c) equilibrium point. Concentration of reactants decreases, and concentration of product increases.

4. The things that remain constant at equilibrium include:
   a. concentration of the reactants and the products.
   b. the rate of forward and reverse reactions.

Answers to Exercise 5.3

The equilibrium constant expression for the following reaction:

1. $NH_3(g) + CH_3COOH(aq) \rightleftharpoons NH^+_4(aq) + CH_3COO^-(aq)$

$$K_c = \frac{[NH_4^+][CH_3COO^-]}{[NH_3][CH_3COOH]}$$

2. $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$

$$K_c = \frac{[H^+][F^-]}{[HF]}$$

3. $2NO(g) + Br_2(g) \rightleftharpoons 2NOBr(g)$

$$K_c = \frac{[NOBr]^2}{[NO]^2\,[Br_2]}$$

4. $2HF(g) \rightleftharpoons H_2(g) + F_2(g)$

$$K_c = \frac{[H_2][F_2]}{[HF]^2}$$

Answers to Exercise 5.4

1. $2NH_3(g) \rightleftharpoons 3H_2(g) + N_2(g)$

From the equation, we see that 2 moles of ammonia produces 3 moles of hydrogen and 1 mole of nitrogen if consumed totally. Here, the amount converted to product is 80.0% of the 5 mole of ammonia (that is 4 mole). At equilibrium, therefore, 1 mole ammonia, 6 mole hydrogen and 2 mole nitrogen are left in the reaction mixture.

$$K_c = \frac{[H_2]^3[N_2]}{[NH_3]^2}$$

$$K_c = \frac{(6/5\text{ mol/L})^3(2/5\text{ mol/L})}{(1/5\text{ mol/L})^2} = 17.3$$

We follow similar procedure for problems 2, 3, and 4 and obtain

2. $K_c = 5.6 \times 10^{-4}$
3. $K_c = 0.0169$
4. $K_c = 0.13$ and $K_p = K_c(RT)^{\Delta n}$. Since $\Delta n = 2\text{-}1=1$, therefore $K_p = K_cRT$.

Answers to Exercise 5.5

1. $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$ $K_c = 1.4 \times 10^{12}$

$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$ $K_c = 5.5 \times 10^{-1}$

You are required to multiply the second equation by 2 and add them to get

$C(s) + CO_2(g) + 2Cl_2(g) \rightleftharpoons 2COCl_2(g)$

Note that 5.5×10^{-1} is multiplied by 2 to give 1.1.

Then the rate constants are multiplied as $1.1 \times 1.4 \times 10^{12}$ to give 1.54×10^{12}.

$$K_c = \frac{[COCl_2]^2}{[CO_2][Cl_2]^2} = 1.54 \times 10^{12}$$

2. Taking $\Delta n = 1$, and using the formula $K_p = K_c(RT)^{\Delta n}$, $K_c = 1.94 \times 10^{-3}$
3. a. $K_c = [O_2]$ b. $K_c = [Ni(CO)_4]/[CO]^4$ c. $K_c = [NOBr]^2/[NO]^2[Br_2]$
4. a. $K_p = 1.4 \times 10^{-5}$ b. $K_p = 266$
5. For the reversible reaction:

$N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $K_p = 0.113$ at 25°C

$$K_p = K_c(RT)^{\Delta n}$$

$$\Rightarrow K_c = \frac{K_p}{(RT)^{\Delta n}} = \frac{0.113}{(8.314\text{J K}^{-1}\text{mol}^{-1} \times 298\text{ K})^{2-1}} = 4.56 \times 10^{-5}$$

the numerical value of K_c at 25°C for the reaction,

$$\tfrac{1}{2}N_2O_4(g) \rightleftharpoons NO_2(g)$$

is half of Kc value for the reaction

$$N_2O_4(g) \rightleftharpoons 2NO_2(g).$$

So, $K_c = \tfrac{1}{2}\,(4.56 \times 10^{-6}) = 2.28 \times 10^{-6}$

Answers to Exercise 5.6

1. For the reaction

$$N_2O_4 \rightleftharpoons 2NO_2$$

the equilibrium constant expression is written as

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = 0.36$$

Initial concentration of N_2O_2 is 0.25 mol/1.5 L, which is 0.167 M.

$$\text{Therefore, } \frac{[NO_2]^2}{[N_2O_4]} = \frac{(2x)^2}{0.167 - x}$$

By solving the equation, be sure you get $[NO_2] = 0.17$ M, $[N_2O_4] = 0.085$ M

2. $2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$ $\qquad K_c = 0.0169$

$$K_c = \frac{[NO]^2[Br_2]}{[NOBr]^2} = \frac{(1.05\times10^{-2}\text{ M})^2(5.24\times10^{-3}\text{ M})}{[NOBr]^2}$$

$$[NOBr]^2 = \frac{(1.05\times10^{-2}\text{ M})^2(5.24\times10^{-3}\text{ M})}{K_c} = \frac{(1.05\times10^{-2}\text{ M})^2(5.24\times10^{-3}\text{ M})}{0.0169}$$

$$[NOBr] = 5.85 \times 10^{-3}\text{ M}$$

3. a. No b. Reverse

4. The concentration table may be constructed as follows:

Equation:	$H_2(g)$	+	$I_2(g)$ $\rightleftharpoons$	$2HI(g)$
Initial(M)	0.33		0.33	0.0
Change(M)	$-x$		$-x$	$+2x$
Equil. (M)	$0.33 - x$		$0.33 - x$	$2x$

The equilibrium constant expression becomes:

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(2x)^2}{(0.33\text{-}x)(0.33\text{-}x)} = 64$$

After solving for x, we get:

$[HI] = 0.532$ M, $[H_2] = [I_2] = 6.7 \times 10^{-2}$ M

5. a. No b. Forward

Answers to Exercise 5.7

1. Ammonia, NH_3, is manufactured by Haber process.
2. The raw materials used in Haber process are natural gas (which is mostly methane) for making hydrogen, and air for making nitrogen.
3. The role of iron in Haber process is to catalyze the reaction between H_2 and N_2 to form NH_3.
4. The reactions made to run faster in the Haber process under optimum conditions of temperature (400°C) and pressure (150–300 atm) and using iron as a catalyst.
5. A flow scheme of the Haber process for manufacturing ammonia is given in the text.
6. The mixture of H_2SO_4 and free SO_3 is known as fuming sulphuric acid or oleum.
7. The raw materials for making SO_2 in the contact process include sulphur or sulphide ores like pyrite, air and concentrated sulphuric acid.
8. SO_3 cannot be simply added to water to make sulfuric acid because the reaction is so uncontrollable that it creates a fog of sulfuric acid. Instead, the sulfur trioxide is first dissolved in concentrated sulphuric acid:

   $H_2SO_4(l) + SO_3(g) \rightarrow H_2S_2O_7(l)$

The product is known as fuming sulphuric acid or oleum. This can then be reacted safely with water to produce concentrated sulphuric acid twice as much as you originally used to make the fuming sulphuric acid.

$H_2S_2O_7(l) + H_2O(l) \rightarrow 2H_2SO_4(l)$

5.2 PHASE EQUILIBRIUM

Periods Allotted: 4 periods

Competencies

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

- *explain how the interplay between kinetic and potential energy underlies the properties of the three states of matter and their phase changes;*
- *explain the process involved, both within a phase and through a phase change, when heat is added or removed from a pure substance;*
- *explain the meaning of vapour pressure and how phase changes are dynamic equilibrium processes;*
- *explain the relationship between vapour pressure and boiling point;*
- *describe how a phase diagram shows a phase of a substance at different conditions of pressure and temperature;*
- *use a phase diagram, determine melting point, boiling point, critical temperature, critical pressure and triple point of a substance.*

Forward Planning

Be advised to read and understand the notes included in the students' text. Refer to other reference books to broaden your understanding of the concepts presented here. Prepare appropriate examples to elaborate the concepts of phase equilibrium. Be sure that the materials required for the activities suggested are ready. You may give a reading homework to your students. Hence, decide on an appropriate time to give the assignment before the class. To motivate their reading, you may set some questions to be answered from the text.

Teaching Aids

No experiment is suggested for this section. Prepare diagrams prepared on flip charts for demonstration of phase diagrams.

Subject Matter Presentation

For the delivery of the lessons in this topic, active learning methodologies such as group discussion, brain storming and gapped lectures using diagrams are generally suggested. This section includes presentation of the phase equilibrium concept in terms of phase, phase rule and phase diagram.

Forces of Attraction, Kinetic Energy and States of Matter

You may use group discussion, brain storming and gapped lectures as suggested active learning methodologies for the delivery of the lessons in this sub-topic. Introduce the phrase phase equilibrium and compare it with chemical equilibrium. *Activity 5.10* will help you to involve your students in discussion. You are required to help the students in forming groups, clarify the procedures followed in the activity where required and provide them the equipment and chemicals required for the activity. Encourage them to think independently and share their ideas. After listening to the presentations of some of the group representatives, continue your lecture to harmonize students' idea with the facts in the text. The following answer helps you to harmonize the ideas suggested by the students.

You will observe that the substances in both the test tubes are in the liquid state. However, in test tube A, only one layer is present, as water and ethanol are completely miscible. In test tube B, two layers of liquids are present, as water and chloroform are immiscible.

Common Terms: Phase, Component and Degree of Freedom

You may apply suggested active learning methodologies such as group discussion, brain storming and gapped lectures for the delivery of the lessons in this sub-topic. Give a reading assignment ahead of time with some key questions. Use the results to identify the area of misunderstanding of the students. Hence, discussion in class is aimed to clear the doubts that the students have. You are required to focus on the basic terms given. Since this topic is difficult, limit its coverage to introducing the terms: phase, phase rule, degree of freedom, etc. Use the example given to stabilize students' understanding.

Phase Rule

For the delivery of the lessons in this sub-topic, you may apply suggested methodologies such as group discussion, brain storming and gapped lectures. You are required to state phase rule and explain it, using the examples that follow. *Activity 5.12* is used to discuss the next topic. You are required to help the students in forming groups, clarify the procedures followed in the activity where required and provide them the equipment and ice cubes for the activity. Encourage them to think independently and share their ideas. The following answer helps you to harmonize the ideas suggested by the students.

When ice is heated under atmospheric pressure, its temperature starts increasing until

0°C is reached, and then the ice starts melting. At this point, when melting is taking place, the temperature remains constant. Both ice and liquid water coexist in dynamic equilibrium. On further heating, more and more ice melts, and ultimately only liquid water remains. When liquid water is further heated, its temperature increases until 100°C is reached, when water starts boiling. On further heating, the water boils away completely to form water vapors.

You may use a mini-lecture to involve the students in questions and answers to discuss the remaining part of the lesson.

Phase Diagram

Study the content and organize the lesson to be covered in one period. Focus on the basic terms given. Let the students also read their textbook ahead of time. Use diagrams prepared ahead of time to make your explanation easy and to save time.

You are required to use a diagram for one component system to introduce the concept. Let students discuss the meaning of Figure 5.6, using the text that follows. During your tour around the groups, find out the problems faced by students in understanding the diagram. In your final presentation, focus your explanation on resolving students' misunderstandings and doubts.

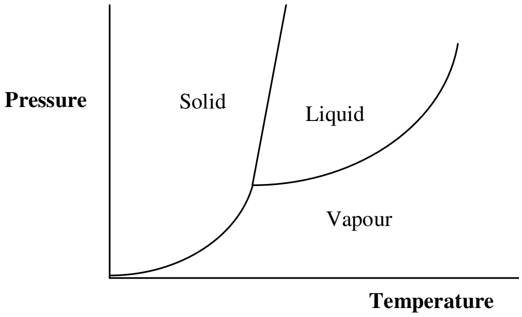
Additional Notes

At its simplest, a phase can be just another term for solid, liquid or gas. If you have some ice floating in water, you have a solid phase present and a liquid phase. If there is air above the mixture, then that is another phase.

But, the term can be used more generally. For example, oil floating on water also consists of two phases - in this case, two liquid phases. If the oil and water are contained in a bucket, then the solid bucket is yet another phase. In fact, there might be more than one solid phase if the handle is attached separately to the bucket rather than moulded as a part of the bucket.

We can recognise the presence of the different phases because there is an obvious boundary between them - a boundary between the solid ice and the liquid water, for example, or the boundary between the two liquids.

A phase diagram lets us work out exactly what phases are present at any given temperature and pressure. In the cases we'll be looking at on this page, the phases will simply be the solid, liquid or vapour (gas) states of a pure substance. The phase diagram for a typical pure substance is represented in the figure below.



These diagrams (including this one) are nearly always drawn in a highly distorted way in order to see what is going on more easily. There are usually two major distortions. If you look at the diagram, you will see that there are three lines, three areas which are marked "solid", "liquid" and "vapour", and two special points marked "C" and "T".

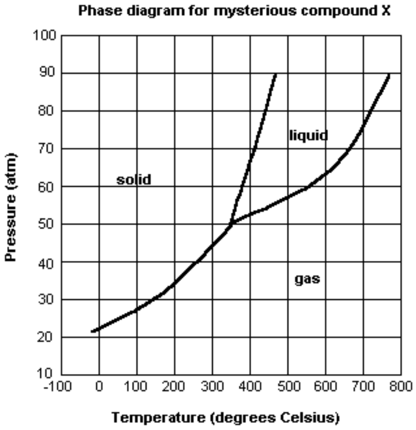
Assessment

Assess each student's work to determine whether the students have achieved minimum required level or not. Give Exercise 5.8 as class work to assess whether the students have achieved minimum required level or not. In Exercise 5.9, let students write their answer independently in their note books. Let them exchange ideas with oneanother and assess themselves. As you give the answers, students will enjoy correcting themselves.

You may use Exercise 5.10 to assess your students. You are required to assess students' active involvement in each of the activities. Your student-performance record list is helpful in all your assessments. Appreciate and give extra activities to students working above the minimum requirement level. You are required to assist those working below the minimum requirement level.

Additional Questions

For each of the questions on this work sheet, refer to the phase diagram for mysterious compound X.



*1. What is the critical temperature of compound X?

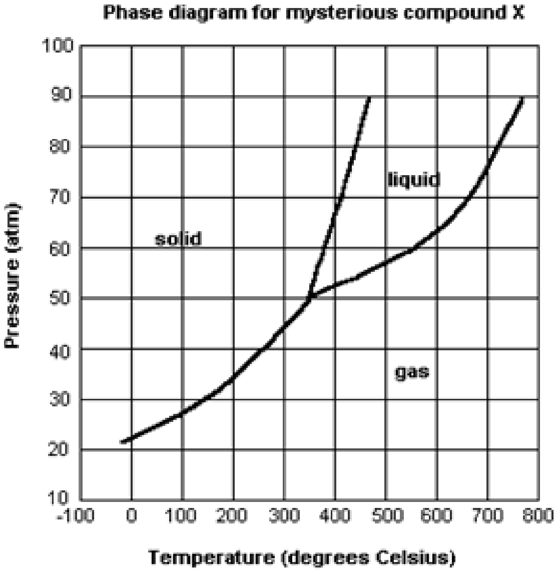
*2. If you were to have a bottle containing compound X in your closet, what phase would it most likely be in?

*3. At what temperature and pressure will all three phases coexist?

4. If we have a bottle of compound X at a pressure of 45 atm and temperature of 100^0 C, what will happen if we raise the temperature to 400^0C?

5. Why can't compound X be boiled at a temperature of 200^0C?
6. If we wanted to, could we drink compound X?

Answers to Additional Questions



For each of the questions on this work sheet, refer to the phase diagram for mysterious compound X.

*1. The critical temperature of compound X is ~770^0 C.

*2. Extrapolating from this diagram, it's most likely a gas.

*3. The temperature and pressure at which all three phases coexist are 350^0C and ~51 atm, respectively.

4. If we raise the temperature to 400^0C, it will sublime.
5. It does not form a liquid at this temperature. It only exists as a liquid at temperatures above 350^0C.
6. No. At the temperatures and pressures that it forms a liquid, you would probably die.

Answers to Exercise 5.8

a. NH_4Cl solid is placed in an evacuated chamber. After a while, some ammonia and HCl appear in the gas phase above it. The number of components, c, is equal to 1 because there is an equilibrium relationship and the number of moles of NH_3 and HCl must be equal.

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

The number of phase, *p*, is equal to 2, which indicates solid and gas phases.

b. The number of phases is one, and the number of components is also one in a system that consists of $COCl_2$, CO and Cl_2 at equilibrium with no excess CO or Cl_2.

c. In the system that consists of $COCl_2(g)$, CO(g), and $Cl_2(g)$ at equilibrium, with excess CO(g) added (i.e., the concentrations of CO and Cl_2 are unequal), the number of phases is one and the number of components is two.

$$COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g)$$

Answers to Exercise 5.9

1. The degree of freedom, F for an equilibrium gas phase containing N_2O_4 and NO_2.
   $N_2O_4(g) \rightleftharpoons NO_2(g)$
   $F = C - P + 2 = 1 - 1 + 2 = 2$ (bivariant system)
2. The degree of freedom, F for phosgene, $COCl_2$, CO and Cl_2 (with CO and Cl_2 derived from decomposition exclusively).
   $COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g)$
   $F = C - P + 2 = 1 - 1 + 2 = 2$ (bivariant system)
3. The degree of freedom, F for phosgene, $COCl_2$, CO and Cl_2, with excess CO added.

   $COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g)$

   $F = C - P + 2 = 2 - 1 + 2 = 3$ (trivariant system)
4. The degree of freedom, F for NH_4Cl, initially in an evacuated chamber, with NH_3 and HCl present from decomposition.
   $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$

   $F = C - P + 2 = 1 - 2 + 2 = 1$ (univariant system)
5. The degree of freedom, F for water, water vapor and ice in equilibrium.
   $H_2O(s) \rightleftharpoons H_2O(l) \rightleftharpoons H_2O(g)$

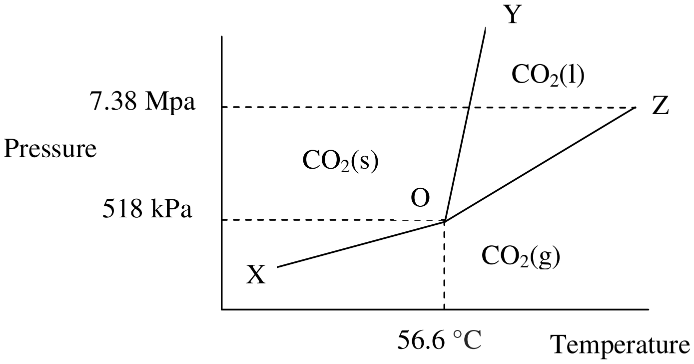
   $F = C - P + 2 = 1 - 3 + 2 = 0$ (invariant system)

Answers to Exercise 5.10

1. Under a pressure of 1 atm, the maximum temperature at which the water can exist in liquid phase is 100 °C.
2. When ice is compressed at 0°C, it will change to water vapor.
3. The point at which ice, liquid water and water vapor can coexist at equilibrium is called triple point. The degree of freedom at this point is calculated as follows:
   $F = C - P + 2 = 1 - 3 + 2 = 0$ (invariant system)

It refers to a triple point on the water phase diagram.

4. The triple point of CO_2 is 518 kPa at 56.6°C. Its critical point lies at 7.38 MPa at 31.1°C. The sublimation curve, fusion curve and vaporization curve have positive slope. Sketch the phase diagram of CO_2 and label it.



Answers to Review Exercises on Unit 5

Part I: Multiple Choice Questions

1. C	2. B	3. C	4. A
5. C	6. A	7. D	8. B
9. C	10. B		

Part II: Answer the following questions

11. For the reaction,

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g) \qquad \Delta H = -196\text{ k J mol}^{-1}$$

a. removal of SO_2 shifts the equilibrium to the left direction.

b. addition of noble gas (argon) at constant pressure does not affect the equilibrium.

c. decreasing the volume of the system shifts the equilibrium to the right direction.

d. increasing the temperature of the system shifts the equilibrium to the left direction.

e. addition of a catalyst has no effect on the equilibrium.

12. Balance the following equations and write the equilibrium constant expressions, in terms of *K*c and *K*p.

a. $2NH_3(g) + 3/2O_2(g) \rightleftharpoons N_2(g) + 3H_2O(g)$

$$K_c = \frac{[N_2][H_2O]^3}{[NH_3]^2[O_2]^{3/2}} \text{ and } K_p = \frac{(p_{N_2})(p_{H_2O})^3}{(p_{NH_3})^2(p_{O_2})^{3/2}}$$

b. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} \text{ and } K_p = \frac{(p_{NH_3})^2}{(p_{N_2})(p_{H_2})^3}$$

c. $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

$K_c = [CO_2]$ and $K_p = p_{CO_2}$

13. For the reaction at 200°C

$$2A(g) + B(g) \rightleftharpoons 3C(g)$$

the equilibrium constant is 3.0. Given the following information,

Species	*Concentration*
[A]	2.0 M
[B]	3.0 M
[C]	2.0 M

Predict the direction in which the reaction should proceed to reach equilibrium.

Solution:

First we need to calculate Q_c

$$Q_c = \frac{[C]^3}{[A]^2[B]} = \frac{(2.0\ M)^3}{(2.0\ M)^2(3.0\ M)} = \frac{8.0\ M^3}{12.0\ M^3} = 0.67$$

Since Q_c (0.67) < K_c (3.0), the reaction should proceed to the right to reach equilibrium.

14. Given the reaction

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$

what is the concentration of CO in equilibrium at 25°C in a sample of gas originally containing 1.00 mol L^{-1} of CO_2?

For the dissociation of CO_2 at 25°C, $Kc = 2.96 \times 10^{-92}$.

Solution:

$$K_c = \frac{[CO]^2[O_2]}{[CO_2]^2}$$

$$2.96 \times 10^{-92} = \frac{(2x)^2(x)}{(1.0-2x)^2}$$

$$2.96 \times 10^{-92} \cong \frac{(2x)^2(x)}{(1.0)^2} \Rightarrow 2.96 \times 10^{-92} \cong 2x^3 \Rightarrow x^3 = 1.48 \times 10^{-92} \Rightarrow x = 2.45 \times 10^{-31}$$

Therefore, the concentration of CO in equilibrium,

$[CO] = 2x = 2(2.45 \times 10^{-31})\ M = 4.91 \times 10^{-31}\ M.$

15. The number of degrees of freedom for a system of a saturated solution of Na_2SO_4 with an excess of solid present in equilibrium with its vapor in a closed vessel is calculated as:
    $F = C - P + 2 = 1 - 2 + 2 = 1$
    The independent variables are temperature and pressure.
16. When the pressure remains constant at 1 atm but the temperature changes from 10°C to 75°C, water remains in the liquid state. However, it changes to water vapor only when the temperature reaches 100 °C.
17. 5.0 moles of ammonia were introduced into a 5.0 L reaction chamber in which it partially dissociated at high temperatures.

    $2NH_3(g) \rightleftharpoons 3H_2(g) + N_2(g)$

At equilibrium at a particular temperature, 80.0 % of the ammonia had reacted.

Calculate *K*c for the reaction.
Solution:

	$2NH_3(g)$	$\rightleftharpoons 3H_2(g)$	$+$	$N_2(g)$
Initial (M)	1.0	0		0
Change (M)	$-2x$	$+3x$		$+x$
Equilibrium (M)	$1.0 - 2x$	$3x$		x

Now, it is given that 80.0 % of the ammonia had reacted which is

$2x = 0.8$ M and $x = 0.4$ M.

So, $[NH_3] = (1.0 - 2x)\ M = (1.0 - 0.8)\ M = 0.2\ M$, $[H_2] = 3x = (3 \times 0.4)\ M = 1.2\ M$, and $[N_2] = x = 0.4\ M$.

Substituting in the K_c expression, we can solve for K_c as follows:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(0.2\ M)^2}{(0.4\ M)(1.2\ M]^3} = \frac{0.4\ M^2}{(0.4\ M)(1.2\ M]^3} = \frac{0.4\ M^2}{0.6912\ M^4} = 0.58\ M^{-2}$$

18. 1.25 mol NOCl was placed in a 2.50 L reaction chamber at 427°C. This gives 1.25 mol/2.50 L = 0.5 mol/L. After equilibrium was reached, 1.10 moles of NOCl remained, which is 1.10 mol/2.50 L = 0.44 mol/L. The equilibrium constant, *K*c, for the reaction is calculates as follows:

Solution:

	$2NOCl(g) \rightleftharpoons$	$2NO(g)$ +	$Cl_2(g)$
Initial (M)	0.5	0	0
Change (M)	$-2x$	$+2x$	$+x$
Equilibrium (M)	$0.5 - 2x$	$2x$	x

Now, it is given that 0.44 mol/L of NOCl remained at equilibrium, which is

$[NOCl] = (0.5 - 2x)$ mol/L = 0.44 mol/L.

Solving for x, we get

$x = (0.5 - 0.44)$ mol/L = 0.06 mol/L

So, $[NO] = 2x = 2(0.06)$ mol/L = 0.12 mol/L, $[Cl_2] = x = 0.06$ mol/L.

Substituting in the K_c expression, we can solve for K_c as follows:

$$K_c = \frac{[NO]^2[Cl_2]}{[NOCl]^2} = \frac{(0.12\text{ mol/L})^2(0.06\text{ mol/L]}}{(0.44\text{ mol/L})^2} = 1.96 \times 10^{-3}\text{ mol/L}$$

19. A sample of nitrosyl bromide, NOBr, was heated to 100°C in a 10.0 L container in order to partially decompose it.

$$2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$$

At equilibrium, the container was found to contain 0.0585 mole of NOBr, 0.105 mole of NO, and 0.0524 mole of Br_2. Calculate the value of Kc.

Solution:

[NOBr] = 0.0585 mol/10.0 L = 0.00585 M, [NO] = 0.105 mol/ 10.0 L = 0.0105 M, and $[Br_2]$ = 0.0524 mol/10.0 L = 0.00524 M.

Substituting in the K_c expression, we can solve for K_c as follows:

$$K_c = \frac{[NO]^2[Br_2]}{[NOCl]^2} = \frac{(0.0105\text{ M})^2(0.00525\text{ M]}}{(0.00585\text{ M})^2} = \frac{(0.0105\text{ M})^2(0.00525\text{ M]}}{(0.00585\text{ M})^2} = 0.90\text{ M}$$

20. The brown gas NO_2 and the colorless gas N_2O_4 exist in equilibrium.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

0.625 mole of N_2O_4 was introduced into a 5.00 L vessel and was allowed to decompose until it reached equilibrium with NO_2. The concentration of N_2O_4 at equilibrium was 0.0750 M. Calculate Kc for the reaction.

Solution:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Initial (M)	0.125	0
Change (M)	$-x$	$+2x$
Equilibrium (M)	$0.125 - x$	$2x$

It is given that the concentration of N_2O_4 at equilibrium was 0.0750 M

$$[N_2O_4] = 0.125 - x = 0.0750 \text{ M}$$

Solving for x, we get

$$x = (0.125 - 0.0750) \text{ M} = 0.05 \text{ M}.$$

So, $[NO_2] = 2x = 2(0.05\text{M}) = 0.1$ M.

Substituting in the K_c expression, we can solve for K_c as follows:

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.1 \text{ M})^2}{(0.0750 \text{ M})} = 0.13 \text{ M}$$

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UNIT 6

CARBOXYLIC ACIDS, ESTERS, FATS AND OILS

UNIT OVERVIEW Total Periods allotted: 20 periods

Dear Teacher! This unit deals with a narrow range of oxygen-containing hydrocarbons that are abundant in nature, especially in living organisms. It mainly deals with the chemistry of carboxylic acids and its derivatives, known as esters, as well as fats and oils.

This unit is structured into three sections as briefly stated below. The first section (6.1) focuses on some general facts about the carboxylic acids. The carboxylic acid functional group is called the *carboxyl group* (–COOH), which consists of the *carbonyl group* (C=O) and the *hydroxyl group* (–OH) bonded to the carbonyl carbon. So, in the functional derivatives of carboxylic acids, either the "H" or the "–OH" group is replaced by some other group. In this section, the structure, nomenclature, preparation, use of carboxylic acids and fatty acids will be presented. Laboratory method of preparing acetic acid by oxidation of ethanol will be demonstrated to the students.

The second section (6.2) is concerned with chemistry of esters. They are chemical compounds derived by reacting an oxoacid (one containing an oxo group, X=O) with a hydroxyl compound such as an alcohol or phenol. Esters are usually derived from an inorganic acid or organic acid in which at least one –OH (hydroxyl) group is replaced by an –O–alkyl (alkoxy) group, and most commonly from carboxylic acids and alcohols. Basically, esters are formed by condensing an acid with an alcohol. The nomenclature, preparation and characteristic reactions of esters will be presented.

Finally, the last section (6.3) deals with the chemistry of fats and oils. In all sections, many worked examples and exercises are included for the purpose of giving the students mor occasions to practice. Please advise your students to take time and study them critically. This unit requires a minimum of *20* periods to complete successfully. We advise you to do the best you can while teaching this unit.

Unit Outcomes

After completing this unit, students will be able to:

- *write the structural formula and IUPAC names of given carboxylic acids and esters;*

- *describe some physical and chemical properties of carboxylic acids and esters;*
- *predict and correctly name the products of organic reaction, including substitution, addition, elimination, esterification, hydrolysis, and oxidation reactions;*
- *carry out activities to prepare a carboxylic acid and an ester;*
- *list some important fatty acids;*
- *test for the carboxylic acid and ester functional groups;*
- *understand the structures, properties, and uses of fats and oils;*
- *describe and explain reactions in soap making (saponification);*
- *describe and explain the cleaning action of soap and detergents;*
- *understand the harmful aspects of detergents on the environment when improperly disposed of;*
- *describe scientific enquiry skills used in this unit: observing, classifying, comparing and contrasting, asking questions, drawing conclusions, applying concepts and problem solving.*

Main Contents

6.1 Carboxylic Acids
6.2 Esters
6.3 Fats and Oils

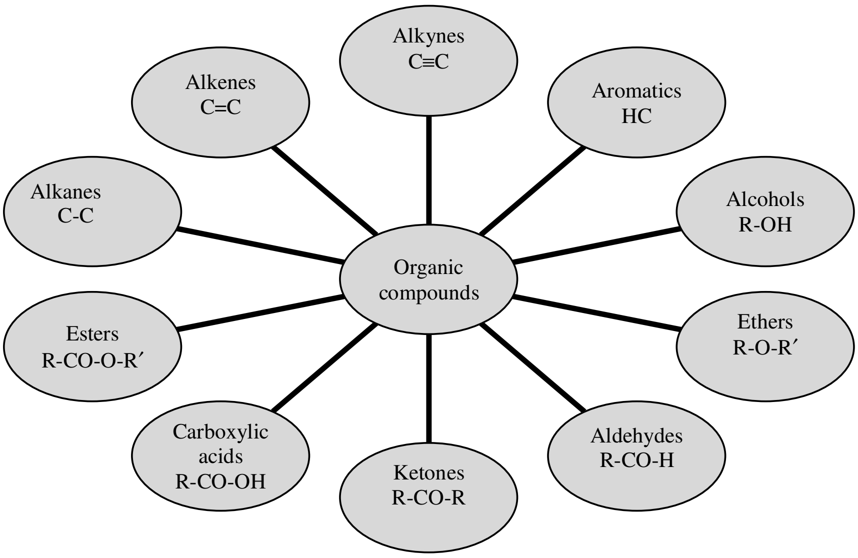
Introduction

Show the students the breakdown of the contents of this unit. Next, you may start with start-up activity. This start-up activity is designed to promote the general thinking of students about carboxylic acids, esters, fats and oils. This activity may be done outside the classroom, taking students to a garden area in the school compound.

The students should collect some common fresh fruits like banana, mango, grapes, apple, pineapple, etc. and some flowers from the garden. They should observe the smell of all the fruits and flowers which they collected. They should also taste the fruits and observe the sourness in each fruit.

Next, guide the students in sharing their ideas with the rest of the class. The students may report to the class that fruits and flowers may have good or bad types of odor. Similarly, some fruits may have sour or sweet tastes. These observations may depend on the difference in the nature of specific chemicals in the fruits or flowers. They are responsible for the odor and taste.

Now, before you start with presentation of carboxylic acids, you may apply concept-mapping methodology and help students begin to understand interrelationships among organic compounds they have studied so far. The students may develop various spider diagram such as the one indicated below.



6.1 CARBOXYLIC ACIDS

Periods Allotted: 8 periods

Competencies

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

- *list common organic acids and mention their sources;*
- *write the general formula of saturated monocarboxylic acids;*
- *write the molecular formulas and names of the first six members of saturated monocarboxylic acids;*
- *give the structural formula for the first four members of saturated monocarboxylic acids;*
- give examples for monocarboxylic, dicarboxylic and tricarboxylic acids;
- *name some branched carboxylic acids;*

- *describe the physical properties of saturated monocarboxylic acids;*
- *explain the general methods of preparation of saturated monocarboxylic acids;*
- *explain the industrial and laboratory preparation of acetic acid;*
- *conduct an experiment to acetic acid in the laboratory;*
- *name and write structural formulas of some fatty acids; and*
- *describe some uses of common carboxylic acids.*

Forward Planning

Set a lesson plan by dividing this topic into separate lessons so as to complete it within 8 consecutive periods. Reading the notes and also practicing the activities as it well as exercises included in the student's textbook before starting presentation is very important. You need to understand the contents well. If you feel that the materials included in the textbook are inadequate, you can also read additional reference materials listed for you at the end of this unit so that you will have access to other examples if students need extra help.

You need to test experiments before you decide to demonstrate them to your students. Collect all the necessary teaching aids which are helpful to the students for demonstration of carboxylic acids

Teaching Aids

Refer to the students' textbook for the apparatus and chemicals required for performing experiment 6.1. Ball-and-stick models and diagrams prepared on flip charts for demonstrating the structure of carboxylic acids are also helpful to the students.

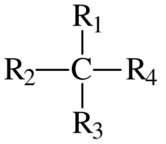
Subject Matter Presentation

Generally, for the delivery of the lessons in this topic, group discussion, brain storming, concept map and presentation are active learning methodologies suggested. This section includes presentation of the chemistry of carboxylic acids. Its major concern is the structure, nomenclature, properties, preparation, and uses of carboxylic acids. Fatty acids are treated in this section. The breakdown of this presentation is briefly presented below.

Structure of Carboxylic Acids

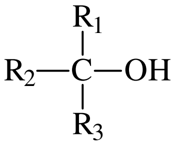
You may use active learning methodologies such as group discussion, brain storming, concept map and presentation for the delivery of the lessons in this sub-topic. After the students discuss the start-up activity, tell them that the first lesson starts with *structure and nomenclature of carboxylic acids*. Now, to link the new topic of discussion to their previous knowledge, help them to discuss *Activity 6.1*. This activity is designed to encourage the students to recall their previous knowledge. Remind the students that they were introduced to hydrocarbons and alcohols in grade 10. You are required to help the students in forming groups, to clarify the questions in the activity, where required, and to facilitate discussions among group members. Appropriate time should also be set for the activity. Let them think independently and share their opinions with their group members. Finally encourage them to report their common understanding. The following answers help you to harmonize the ideas developed through discussion.

Alkanes have the general structure of



$$\begin{array}{ccc} & R_1 & \\ & | & \\ R_2 - & C & - R_4 \\ & | & \\ & R_3 & \end{array}$$

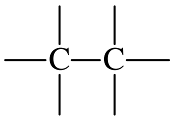
and alcohols have the general structure of



$$\begin{array}{ccc} & R_1 & \\ & | & \\ R_2 - & C & - OH \\ & | & \\ & R_3 & \end{array}$$

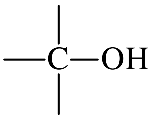
where R_1, R_2, R_3, and R_4 are alkyl groups or hydrogen atom.

The two structures differ in the presence of the hydroxyl group, R–OH, in alcohols and its absence in alkanes. The structure which determines the properties of alkanes is the presence of the C–C bond.



$$\begin{array}{cccc} & | & | & \\ - & C - & C & - \\ & | & | & \end{array}$$

The presence of C–OH determines the properties of alcohols.



$$\begin{array}{ccc} & | & \\ - & C & - OH \\ & | & \end{array}$$

From this activity, we can conclude that properties of a specific organic compound depend on its structure.

Next, continue the discussion by presenting the structure of carboxylic acids with a breakdown of the structure of saturated mono–, di– and tricarboxylic acids. Now, help the students to discuss *Activity 6.2*. This activity is designed to encourage the students to identify some fruits having an acidic taste. The following answers help you to harmonize the ideas developed through discussion.

Type	***Predominant Acids Present***
Lemons	Citric, malic, tartaric and oxalic acids (no isocitric acid)
Orange	Citric, malic and oxalic acids
Bananas	Malic, citric, tartaric and traces of acetic and formic acids
Apples	Malic, quinic, aketoglutaric, oxalacetic, citric, pyruvic, fumaric, lactic and succinic acids
Pineapples	Citric and malic acids.
Avocados	Tartaric acid acids
Salad	Citric and malic acids
Grapefruit	Citric, tartaric, malic and oxalic acids.

Finally, Exercise 6.1 should be given as homework, which you will use part of the continuous assessment.

Common Names of Carboxylic Acids

Here, you may assign students to read and discuss the subject in groups. As you walk around the classroom and facilitate their discussions, try to find out the difficulties students face in understanding the content of the text.

The students should know common names of carboxylic acids. This lesson will have the following breakdown:

- common names of straight chain monocarboxylic acids;
- common names of branched chain and substituted carboxylic acids;
- common names of dicarboxylic acids; and
- common names of aromatic carboxylic acids.

After students finish the above discussion, present an explanation that will resolve any problems the students have. Exercise 6.2 should be given as homework, which will be included in the continuous assessment. You may also assign them the task of studying the next topic. You may give them the following activities to motivate their reading.

1. Write the detailed structure of octanoic acid.
2. Give a common name of

   $CH_3-(CH_2)_3-CH(CH_3)-COOH$

IUPAC Names of Carboxylic Acids

For the delivery of the lessons in this sub-topic, you may use group discussion, brain storming and concept map as active learning methodologies. Start this lesson by assessing the assignment given to students. Let some of the students give their answer to the question given during the last lesson. Before giving feedback, let them form into groups and read the text and discuss together.

The students should discuss

- IUPAC names of straight chain monocarboxylic acids;
- IUPAC names of branched chain and substituted carboxylic acids;
- IUPAC names of dicarboxylic acids; and
- IUPAC names of aromatic carboxylic acids.

After students finish the above discussion, you have to clear out any problems the students have using a mini-lecture. Exercise 6.3 should be given as homework, which will be part of continuous assessment.

Physical Properties of Carboxylic Acids

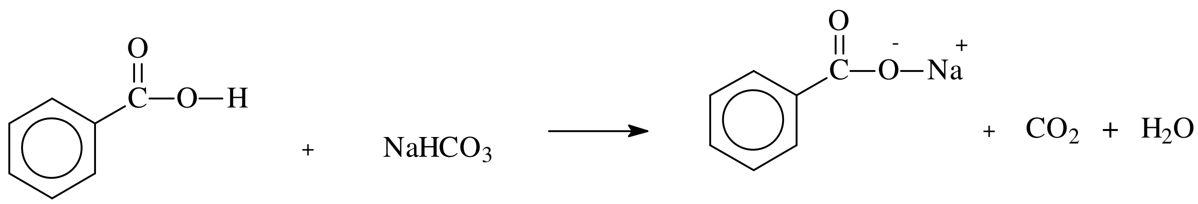
Ask the students to discuss *Activity 6.3* in groups. The purpose of this activity is to recall and use their previous knowledge of physical properties and to use it to explore the physical properties of carboxylic acids. Let the students think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. Harmonize the ideas developed through discussion, based on the following suggested answers.

Alcohols possess higher boiling points than that of alkanes of comparable molar masses. This is because alcohols form intermolecular hydrogen bonding, whereas alkanes do not. Help the students to explore the physical properties of carboxylic acids by looking at the group in a little more detail. You may give them Exercise 6.4 as homework.

Chemical Properties of Carboxylic Acids

You may apply active learning methodologies such as group discussion, brain storming, concept map and presentation for the delivery of the lessons in this sub-topic. This is the fifth lesson of the unit. Ask the students to discuss *Activity 6.4* in groups. This activity is designed to help students recall and use their previous knowledge of chemical properties and to use it to explore the chemical properties of carboxylic acids. You are required to help the students in forming groups, clarify the procedures followed in the activity, where required, and provide them with the equipments and chemicals required for the activity. Encourage them to think independently and share their ideas. The following suggested answers help you to harmonize the ideas suggested by the students.

In this activity, benzoic acid reacts with sodium bicarbonate and gives sodium benzoate, as shown in the equation below.

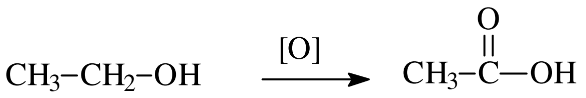


Benzoic acid Sodium bicarbonate Sodium benzoate

The color of the litmus solution before benzoic acid added is red. When benzoic acid is added, drop by drop, the color of litmus solution starts changing and finally is converted to blue. This indicates that the product, sodium benzoate, is formed which gives a basic solution in aqueous solution.

In the main body of discussing the chemical properties of acids, you may let students read through the text and share their understanding with their friends. Ask them whether they can explain the main reactions considered in this section. Then let them perform *Activity 6.5.* This activity is designed to encourage the students to apply their local knowledge about *tej* and *tella.* Let them think independently and share their opinions with their group members. You may propose other questions also related to the souring of the local alcoholic drinks mentioned. (The differences that take place when the container is closed and when it is left open to the atmospheric air). Harmonize the ideas developed through discussion, based on the following suggested answers.

'Tella' or 'tej' turn sour some days after they are made. This happens due to the conversion of beverage alcohol, ethanol, to ethanoic acid.



$$CH_3-CH_2-OH \xrightarrow{[O]} CH_3-\overset{\overset{\displaystyle O}{||}}{C}-OH$$

Ethanol Ethanoic acid

The oxygen comes from atmospheric oxygen. The intermediate compound (ethanol, an aldehyde) can also be mentioned, according to the interest and participation of the students.

Preparation and Uses of Carboxylic Acids

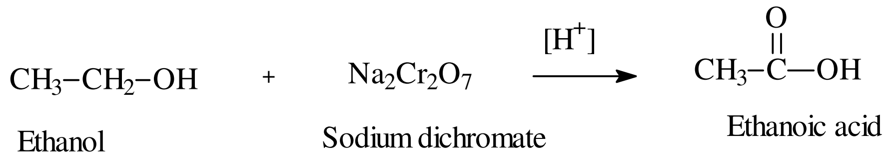
Encourage the students to understand that oxidation is one of the important methods for the preparation of carboxylic acids. The students should emphasize to selected methods of preparing carboxylic acids, which are the oxidation of primary alcohols, and oxidation of alkylbenzenes. The students should also appreciate some of the uses of carboxylic acids. Carboxylic acids are used in industry for making soaps, detergents, and shampoos.

The students should perform Experiment 6.1, which is designed to familiarize students with the laboratory method of preparing acetic acid. Refer to the students' textbook for the detailed procedures as well as the apparatus and chemicals required for performing this experiment.

The predict-observe-explain methodology is helpful in demonstrating this experiment to the students**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

The answers to the questions listed under observation and analysis are given below.

The color of the solution in the flask remains colorless. The smell of the product, the liquid distilled, is different from the smell of ethanol. The role of sodium dichromate is to oxidize ethanol to ethanoic acid as oxidizing agent; it is reduced. The chemical equation for this reaction is:



$$\underset{\text{Ethanol}}{CH_3{-}CH_2{-}OH} + \underset{\text{Sodium dichromate}}{Na_2Cr_2O_7} \xrightarrow{[H^+]} \underset{\text{Ethanoic acid}}{CH_3{-}\overset{\overset{\displaystyle O}{||}}{C}{-}OH}$$

From this activity, students may conclude that acids can be prepared in the laboratory by oxidation of primary alcohols.

Fatty Acids

Give *Activity 6.6* to the students so that they discuss the question and share their ideas with the rest of the class. The purpose of this activity is to encourage the students to express their opinion to their friends regarding what they know about fatty acids. Harmonize the ideas developed through the discussion based on the following suggested answers.

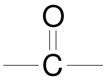
Fatty acids are carboxylic acids with long hydrocarbon chains either with carbon-carbon double bonds (unsaturated fatty acids) or without carbon-carbon double bonds (saturated fatty acids).

After the discussion is over, help the students to:

- know the meaning of a fatty acid;
- give examples of naturally occurring fatty acids; and
- write the structure of simple fatty acids.

Additional Notes

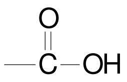
Carboxylic acids are like aldehydes and ketones in that they contain the carbonyl functional group.


$-\overset{\overset{O}{\|}}{C}-$

carbonyl

–OH

hydroxyl


$-\overset{\overset{O}{\|}}{C}-OH$

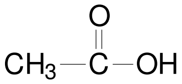
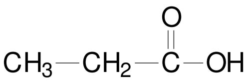
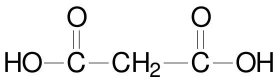
carboxylic acid

Carboxylic acids are the acids found in organic molecules, the carbon based compounds necessary for life. The term "carboxylic" is derived from carbonyl and "hydroxyl", the two structures which make an organic acid. None of these carboxylic acids are classified as strong acids. Remember a strong acid is an acid which fully dissociates in an aqueous environment. This is not to say an organic acid is not able to cause harm. Small concentrations of hydrogen ions given off by organic acids are potent enough to kill most microscopic organisms and cause tissue irritation and/or damage to larger organisms. Acetic acid is the acid found in vinegar. A pickle stored in its brine may not spoil and be edible for years partly because of the organic acid from the vinegar kills the bacteria which would have caused spoilage.

Nomenclature of Carboxylic Acids

1. Count the number of carbons in the longest carbon chain containing the –COOH group.
2. Replace the –e with the suffix –oic acid.
3. Compounds containing multiple -COOH groups do not drop the –e, but add a di- or tri- to the ending – carboxylic acid or add a di- or tri- to the suffix –oic acid.

For example,

Structural Formula	 $CH_3-\overset{\overset{O}{\|}}{C}-OH$	 $CH_3-CH_2-\overset{\overset{O}{\|}}{C}-OH$	 $HO-\overset{\overset{O}{\|}}{C}-CH_2-\overset{\overset{O}{\|}}{C}-OH$
Condensed Structural Formula	CH_3COOH	CH_3CH_2COOH	$HOOCCH_2COOH$
IUPAC	Ethanoic acid	Propanoic acid	Propandioic acid
Common Name	Acetic acid	Propionic acid	Malonic acid

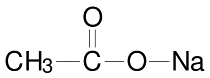
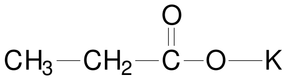
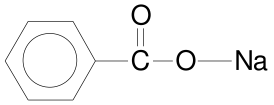
Salts of Carboxylic Acids

In order to create a salt of a carboxylic acid, neutralize an organic acid with a strong base:

a. change the name of the carboxylic acid suffix from –ic acid to –ate, just as you did with the ester.

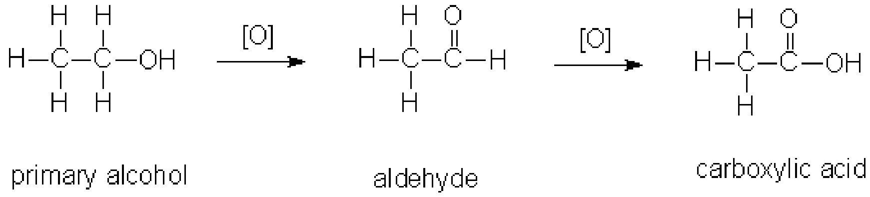
b. the first word will be the name with the cation, probably a metal.

For example,

Structural Formula	 $CH_3-\overset{O}{\overset{\|}{C}}-O-Na$	 $CH_3-CH_2-\overset{O}{\overset{\|}{C}}-O-K$	
Condensed Structural Formula	CH_3COONa	CH_3CH_2COOK	C_6H_5COONa
IUPAC	sodium ethanoate	potassium propanoate	sodium benzoate
Common Name	sodium acetate	-	-

Preparation of Carboxylic Acids

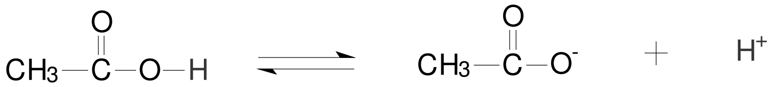
The oxidation of a primary alcohol leads through an aldehyde to a carboxylic acid.



Reactions of Carboxylic Acids

a. Ionization

The first definition of an acid is a proton donator. A carboxylic acid donates the hydrogen attached to the oxygen, fulfilling the definition of an acid.

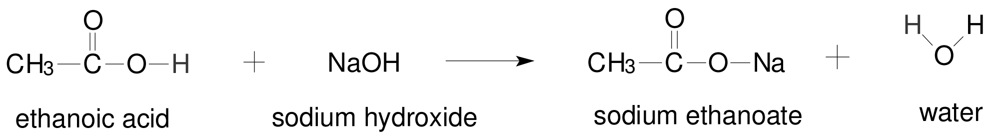


$$CH_3-\overset{O}{\overset{\|}{C}}-O-H \rightleftharpoons CH_3-\overset{O}{\overset{\|}{C}}-O^- + H^+$$

This reaction has a double arrow. The reason is that organic acids are weak acids and form equilibrium in a solution. Only a small percentage of the acids give up their hydrogen. Once the equilibrium has been reached, the concentration of acid, hydrogen ion and anion remain constant. In the above case, we have acetic acid, acetate anion and hydrogen ion.

b. Acid/Base reaction

Just as you can neutralize an inorganic acid, an organic acid can be neutralized. In both cases, the same products are formed, namely a *salt* and *water*. If sodium hydroxide is used to neutralize acetic acid, the following reaction occurs.



$$CH_3-\overset{O}{\overset{\|}{C}}-O-H + NaOH \longrightarrow CH_3-\overset{O}{\overset{\|}{C}}-O-Na + H_2O$$

c. Acid Anhydrides

Two carboxylic acids can be combined to form an acid and anhydride. The term anhydride combines an- and -hydride. The prefix an- means anti, in this form it is describing a loss. The loss is of water, which is where the hydride portion comes to play. So, two carboxylic acids are

$$CH_3-\overset{O}{\overset{\|}{C}}-O-H + H-O-\overset{O}{\overset{\|}{C}}-CH_3 \longrightarrow CH_3-\overset{O}{\overset{\|}{C}}-O-\overset{O}{\overset{\|}{C}}-CH_3 + H_2O$$

Common Sources of Carboxylic Acids

The carboxylic acid function group can be found in all facets of nature. A few of the more common acids follow, formic acid, a chemical defense for some ants, acetic acid, found in vinegar, propanoic acid, found in cheese, butanoic acid, from spoiled butter, citric acid, from citric fruits, lactic acid, from dairy products, palmitic acid, found in palm oil and stearic acid, from beef fat.

Assessment

Assess each student's work to determine whether the students have achieved minimum required level or not. In order to do this, you can use different instruments of evaluation such as class work, homework, quiz, test and examination. Students working at the minimum requirement level will be able to list common organic acids and mention their sources, give the general formula of saturated monocarboxylic acids, give common and IUPAC names of the first six members of saturated monocarboxylic acids, write the molecular formulas and structural formulas for the first four members of saturated monocarboxylic acids, describe the physical and chemical properties of saturated monocarboxylic acids, explain the general methods of preparation of saturated monocarboxylic acids, name and write structural formulas of some fatty acids, and describe some uses of common carboxylic acids.

Praise students working above the minimum requirement level and recognize their achievements. Encourage them to continue working hard and not to become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class. Give extra attention in class and additional lesson time during breaks or at the end of the day.

Additional Questions

Questions 1 to 5 refer to the four carboxylic acids A, B, C and D listed below.

A. ethanoic acid B. ethanedioic acid C. butanoic acid D. decanoic acid

*1. Which of these acids is the least soluble in cold water?

*2. Which of these acids has the lowest boiling point?

*3. Which acid can form two sodium salts?

*4. Which of these acids is called oxalic acid?

*5. Which of these acids, when concentrated, is called ‘glacial’ because in cold weather it crystallises?

6. Which is the best description of the reaction between C_3H_7OH and HCOOH?

A. addition B. hydrolysis C. esterification D. elimination

7. When phenylethanoate is hydrolysed by excess aqueous sodium hydroxide, the products of the reaction are

A. phenol and ethanoic acid B. phenoxide ions and ethanoate ions

C. ethanol and benzoic acid D. phenol and ethanoate ions

8. Which of these organic compounds bubbles when magnesium is added?

A. CH_3CHO B. HCOOH C. CH_3COCH_3 D. HCHO

9. Which of the following compounds would react with the product of its own oxidation product to form a sweet smelling liquid?

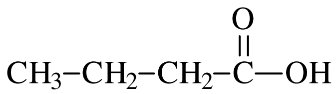
A. propanal B. propanone C. propan-1-ol D. propan-2-ol

Answers to Additional Questions

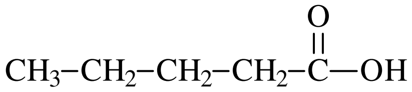
*1. D *2. A *3. B *4. B *5. A 6. C 7. A 8. B 9. C

Answers to Exercise 6.1

1.

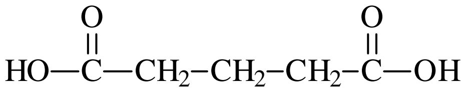
a)


$CH_3{-}CH_2{-}CH_2{-}\overset{\overset{\displaystyle O}{||}}{C}{-}OH$ Butanoic acid

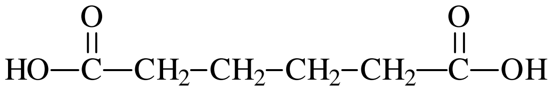
b)


$CH_3{-}CH_2{-}CH_2{-}CH_2{-}\overset{\overset{\displaystyle O}{||}}{C}{-}OH$ Pentanoic acid

2.

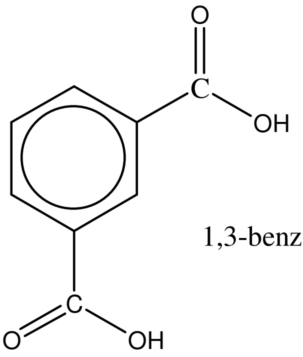
a) 

$HO-\overset{O}{\overset{\|}{C}}-CH_2-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-OH$ Pentanedioic acid

b) 

$HO-\overset{O}{\overset{\|}{C}}-CH_2-CH_2-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-OH$ Hexanedioic acid

c)

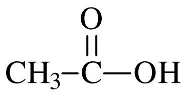


1,3-benzenedicarboxylic acid

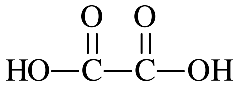
Answers to Exercise 6.2

1. a. α–Bromo–β–Chlorobutyric acid b) Succinic acid

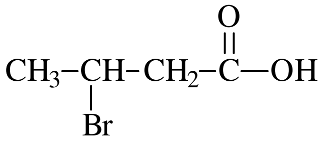
2.

a) 

$CH_3-\overset{O}{\overset{\|}{C}}-OH$ Acetic acid

b) 

$HO-\overset{O}{\overset{\|}{C}}-\overset{O}{\overset{\|}{C}}-OH$ Ethanedioic acid

c) 

$CH_3-\underset{Br}{\underset{|}{CH}}-CH_2-\overset{O}{\overset{\|}{C}}-OH$ β-Bromobutyric acid

Answers to Exercise 6.3

1.

a). 3–Methylbutanoic acid

b). 5–Bromopentanoic acid

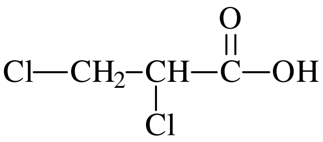
c). 2–Hydroxypropanoic acid

d). 4 – Bromo–3–Chloropentanoic acid

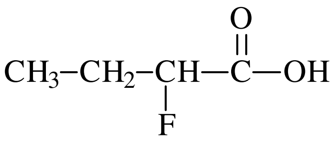
e). 2–Hydroxybenzoic acid

f). 3–Aminobenzoic acid

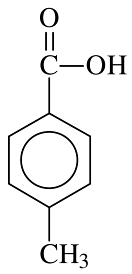
2.

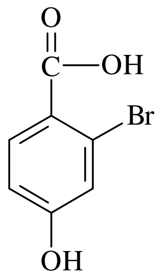
a) 

$$\mathrm{Cl{-}CH_2{-}\underset{\displaystyle Cl}{\underset{|}{CH}}{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}OH}$$

c) 

$$\mathrm{CH_3{-}CH_2{-}\underset{\displaystyle F}{\underset{|}{CH}}{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}OH}$$

b) 

d) 

Answers to Exercise 6.4

1. Decanoic acid > octanoic acid > butanoic acid > propanoic acid.
2. The aliphatic carboxylic acid, which is most soluble in water, is ethanoic acid.
3. The aliphatic carboxylic acid, which has the lowest boiling point, is methanoic acid.
4. C_5H_{12} (Pentane) < $C_4H_{11}OH$ (Butanol) < C_2H_5COOH (Propanoic acid) < $CH_3(CH_2)_4COOH$ (Hexanoic acid).

6.2 ESTERS

Periods Allotted: 6 periods

Competencies

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

- *list common sources of esters;*
- *write the general structural formula of esters;*
- *write the molecular formulas and names of some simple esters;*
- *describe the physical properties of esters;*
- *explain the chemical properties of esters;*
- *explain the general methods of preparation of esters;*
- *describe some uses of common esters.*

Forward Planning

Prepare a daily lesson plan by dividing this topic into separate lessons so as to complete the subject within 6 consecutive periods. Reading the notes and also practicing the activities and the exercises included in the student's textbook before starting presentation is very important. You need to understand the contents well. If you feel that the materials included in the textbook are inadequate, you can read additional reference materials listed for you at the end of this unit so that you will have access to other examples if students need extra help. Collect all the necessary teaching aids which are helpful to the students for the demonstration of esters.

Teaching Aids

No experiment is suggested for this section. But, ball-and-stick models and diagrams prepared on flip charts for demonstrating the structure of esters are helpful to the students.

Subject Matter Presentation

For the delivery of the lessons in this topic, generally you may apply suggested active learning methodologies such as group discussion, investigation and presentation. This section includes the presentation of the chemistry of esters with major emphasis on their sources, structure, nomenclature, properties, preparation, and uses. The breakdown of this presentation is briefly presented below.

Sources and Structure of Esters

Let students perform *Activity 6.7*. This activity is designed to bring chemistry close to the students. The following suggested answers help you to harmonize the ideas developed through discussion.

The students should respond to the question as explained below. The smell of a chemical depends on properties such as the shape, chemical structure or electrical charge of its molecules. Therefore, organic chemists are often taught to identify certain functional groups in an organic compound by its smell while analyzing organic compounds. A fruity smell is commonly associated with an ester, the smell of bitter almonds with ketones and that of rotten eggs with compounds having -SH functional group.

The taste of any food item depends on its chemical composition. A food item is invariably made up of a number of chemical compounds. Different food items, e.g., different kinds of fruits, have different chemical compositions and hence have different tastes. During the process of the chemical substances present in undergo some chemical changes, which transform the structures of some of the molecules, which produce a particular taste.

After harmonizing the discussion on the activity given, let the students read the text and discuss, in groups, the structure of esters. As you tour around, help them to correctly write structure of esters and explain their difference from carboxylic acids.

Nomenclature of Esters

Let the students read the text and discuss, in groups, the nomenclature of esters. You may create a relevant class activity to involve them in group discussion. As you facilitate their discussion, help them to correctly name esters. The students should practice Table 6.7 and do Exercise 6.5 as home work.

Physical Properties of Esters

Let students perform *Activity 6.8*. The purpose of this activity is to encourage the students to recall their knowledge and relate it with physical properties of esters. You are required, to help the students in forming groups, clarify the questions in the activity, where required and facilitate discussion among group members. Let them think independently and share their opinion with their group members. Finally, encourage them to report their common understanding. Harmonize the ideas developed through discussion, based on the following suggested answers.

Esters have lower boiling points compared to carboxylic acids of comparable molecular mass. Esters have lower boiling points compared to alcohols of comparable molecular mass. Let students read the text note and explain the subject to their friends. After summing up the lesson, give them a reading assignment.

Chemical Properties of Esters

Let students perform *Activity 6.9*. This activity is prepared to help the students to recall their knowledge and apply it in the study of chemical properties of esters. You are required to help the students in forming groups, clarify the questions in the activity, where required, and facilitate discussion among group members. Help them to recall the chemical reactions of carboxylic acids. Let them relate this discussion to their prior knowledge. The following suggested answers help you to harmonize the ideas developed through discussion.

When esters are treated with water, they undergo a hydrolysis reaction with water, which yields corresponding carboxylic acids and alcohols. The hydrolysis reaction is speeded up in the presence of a mineral acid, which acts as a catalyst in this reaction. The condensation reaction between the reactants that form esters, gives ester and water molecules.

Preparation of Esters

You may create a relevant class activity to involve them in group discussion. You may relate this lesson with *Experiment 5.1.* Let students recall the reactants, procedures followed and results obtained in the experiment. Relate the lesson also with the chemical properties of carboxylic acids. Finally you may give them *Exercise 6.6* as homework.

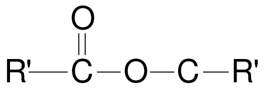
Uses of Esters

You may create a relevant class activity to involve them in group discussion. Let students focus on the industrial uses of esters.

Additional Notes

General formula of carboxylic acids

Esters are represented by the general formula:



$$R'-\overset{\overset{O}{\|}}{C}-O-C-R'$$

ester

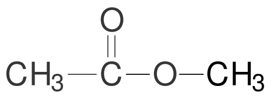
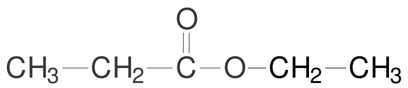
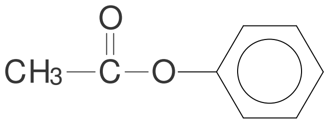
where R and R’ are alkyl groups.

Nomenclature of Esters

An ester is named from its starting materials, the acid and the alcohol.

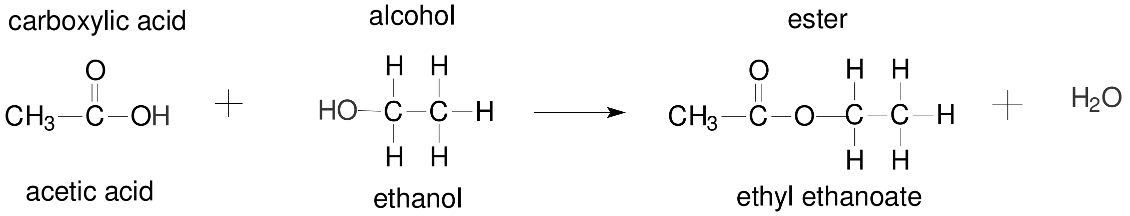
1. the first part of its name comes from the alcohol. In this case, use the side chain abbreviation i.e., methyl, ethyl…etc
2. the second part of its name comes from the carboxylic acid. To end the second part, change the “–ic” of the carboxylic acid to “–ate”.

Examples:

Structural Formula	***Condensed Structural Formula***	***IUPAC***	***Common Name***
 $CH_3-\overset{\overset{O}{\|}}{C}-O-CH_3$	CH_3COOCH_3	Methyl ethanoate	Methyl acetate
 $CH_3-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_2-CH_3$	$CH_3CH_2COOCH_2CH_3$	Ethyl propanoate	Ethyl propyrate
	$CH_3COOC_6H_5$	Phenyl ethanoate	Phenyl acetate

Preparation of esters (esterification)

To create an ester, an alcohol combines with a carboxylic acid in a dehydration reaction, which produces the ester and a water molecule.



Common sources of esters

The most utilized ester is that found in soap. Soap pre-dates history. The production is simple, boil a fat in wood ash. The fat is actually the carboxylic acid, hence the name fatty acid. A fat like stearic acid, in beef fat, would work very well for producing soap. A component of wood ash is potassium hydroxide, which is a strong base. So, your average caveman could easily produce soap by boiling some beef fat combined with ash from the campfire. Other uses include dry cleaner fluids, pharmaceutical drugs, flavorings and scents. Methyl salicylate can be found in wintergreen flavored mints.

Assessment

Assess each student's work to determine whether the students have achieved minimum required level or not. In order to do this, you can use different instruments of evaluation such as class work, homework, quiz, test and examination. Students working at the minimum requirement level will be able to list common sources of esters, write the general structural formula of esters, write the molecular formulas and names of some simple esters, describe the physical and chemical properties of esters, explain the general methods of preparation of esters, and describe some uses of esters.

Praise students working above the minimum requirement level and recognize their achievements. Encourage them to continue working hard and not to become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with rest of the class. Give extra attention in class and additional lesson time during breaks or at the end of the day.

Additional Questions

*1. Name an ester which is responsible for flavor of the following fruits.

a. Methyl butanoate (apple)

b. Methyl methanoate (rum essence)

c. Methyl butanoate (pineapple)

d. Pentyl ethanoate (banana)

e. Octyl butanoate (orange)

2. Complete the table by naming the ester and/or identifying the original alcohol and carboxylic acid it was made from.

Ester	*Alcohol*	*Carboxylic Acid*
$CH_3COOC_2H_5$	(a)	(b)
(c)	CH_3OH	C_2H_5COOH
$HCOOC_3H_7$	(d)	(e)
(f)	C_4H_9OH	$CH_3CH_2CH_2COOH$
$C_3H_7COOC_3H_7$	(g)	(h)

Answers to Additional Questions

*1. a. Methyl butanoate (apple) b. Methyl methanoate (rum essence)

c. Methyl butanoate (pineapple) e. Pentyl ethanoate (banana)

f. Octyl butanoate (orange)

2. Complete the table by naming the ester and/or identifying the original alcohol and carboxylic acid it was made from.

a. HOC_2H_5 b. CH_3COOH

c. $C_2H_5COOCH_3$ d. HOC_3H_7

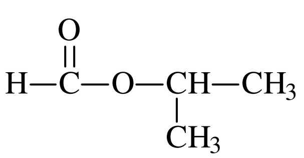
e. $HCOOH$ f. $CH_3CH_2CH_2COOC_4H_9$

g. HOC_3H_7 h. C_3H_7COOH

Answers to Exercise 6.5

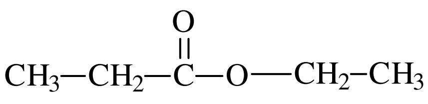
1. a) Propyl methanoate b) Phenyl ethanoate c) Methyl benzoate d) Cyclohexyl propanoate

2) a)



$$H-\overset{\overset{O}{\|}}{C}-O-\underset{\underset{CH_3}{|}}{CH}-CH_3$$

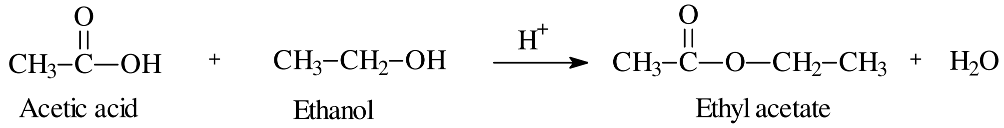
b)



$$CH_3-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_2-CH_3$$

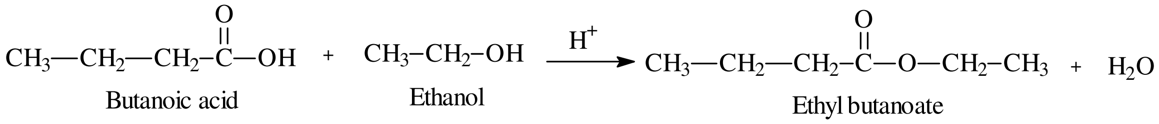
Answers to Exercise 6.6

a)



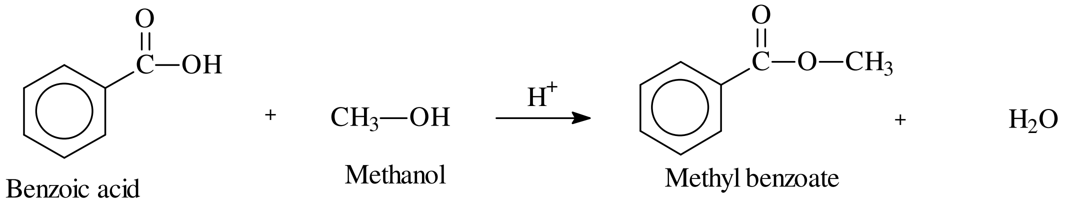
$$\mathrm{CH_3{-}\overset{O}{\overset{\|}{C}}{-}OH + CH_3{-}CH_2{-}OH \xrightarrow{H^+} CH_3{-}\overset{O}{\overset{\|}{C}}{-}O{-}CH_2{-}CH_3 + H_2O}$$

b)

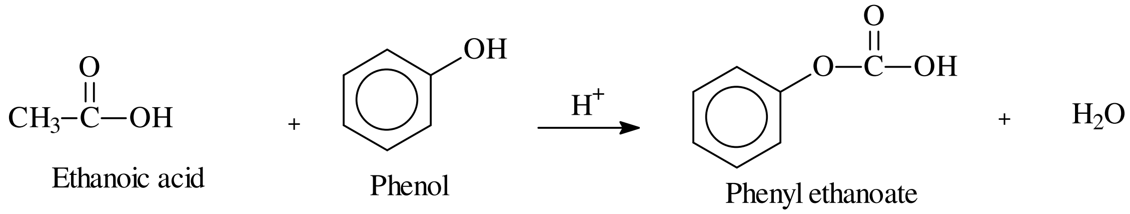


$$\mathrm{CH_3{-}CH_2{-}CH_2{-}\overset{O}{\overset{\|}{C}}{-}OH + CH_3{-}CH_2{-}OH \xrightarrow{H^+} CH_3{-}CH_2{-}CH_2{-}\overset{O}{\overset{\|}{C}}{-}O{-}CH_2{-}CH_3 + H_2O}$$

c)



d)



6.3 FATS AND OILS

Periods Allotted: 6 periods

Competencies

After studying this topic, the students will be able to:

- *define fats and oils;*
- *write the general structural formula for fats and oils;*
- *write the structures of some common trigycerides;*
- *describe physical properties of fats and oils;*
- *explain hardening of oils (process of converting oils to hard fats);*
- *explain rancidity;*
- *define soap and detergent;*
- *explain saponification;*
- *prepare soap;*
- *explain the cleaning action of soaps.*

Forward Planning

Set a lesson plan by dividing this topic into separate lessons so as to complete it within 6 consecutive periods. Reading the notes, and also practicing the activities as well as exercises included in the student's textbook before starting presentation is very important. You need to understand the contents well. If you feel that the materials included in the textbook are inadequate, you can also read additional reference materials listed for you at the end of this unit so that you will have access to other examples if students need extra help.

You need to test experiments before you decide to demonstrate them to your students. Collect all the necessary teaching aids which are helpful to the students for the demonstration of fats and oils. Materials required for Experiment 6.2 should be ready ahead of time. You are also required to attempt the experiment alone for laboratory practice. This section is the end of the unit. Hence, you need to plan to assign your students the task of performing the review exercises.

Teaching Aids

Refer to the students' textbook for the apparatus and chemicals required for performing experiment 6.2. Again, using ball-and-stick models for demonstrating the structure of fats and oils is helpful to the students.

Subject Matter Presentation

Here, suggested active learning methodologies such as enquiry, group discussion and experimentation and peer teaching may be used for the delivery of the lessons in this topic. This section includes the presentation of the chemistry of fats and oils, with major emphasis on their sources, structure, and properties. The concepts of the hardening of oil, rancidity, soaps, and detergents are also treated in this section. The breakdown of this presentation is briefly presented below.

Structure of Fats and Oils

Let students perform *Activity 6.10*, which is designed to help students associate the sources and the physical states of fats and oils. The following suggested answers help you to harmonize the ideas developed through discussion.

Fats are a class of energy-rich organic compounds occurring widely in organic tissue, especially in the adipose tissue of animals and in the seeds, nuts, and fruits of plants. Fats differ from oil in that they are solid or semisolid at room temperature, while oils are liquid. Students should collect the samples of butter, lard, tallow, peanut oil, soyabean oil and olive oil. Peanut oil, soyabean oil and olive oil are vegetable in origin and butter, lard, and tallow are animal in origin. There is a relationship between the source and their physical state. Plant-origin fats are liquids, and animal-origin fats are solids.

Physical Properties of Fat and Oils

Activity 6.11 helps students to differentiate between vegetable oils and mineral oils. You may also create other activities which can motivate and involve students in discussing the topic. Harmonize the ideas developed through discussion, based on the following suggested answers.

Mineral oils edible vegetable oils and are entirely different; mineral oils (petroleum) are long chains of alkanes, whereas vegetable oils are long chains of fatty acid. Mineral oil is not suitable to our bodies and should not be used as food.

Hardening of Oils

Activity 6.12 helps students categorize fatty acids as saturated and unsaturated. You are required to help the students in forming groups, clarify the questions in the activity, where required, and facilitate discussion among group members. Let them think independently and share their opinions with group members. Finally, encourage them to report their common understanding. The following suggested answers help you to complement the ideas developed through discussion.

Unsaturated fatty acids are oleic, linolic, and linolenic acids. Myristic, lauric, palmitic, and stearic acid are saturated fatty acids. Unsaturated fatty acids can be converted to saturated fatty acids by adding hydrogen (hydrogenation) under high pressure and in presence of catalyst.

Rancidity

Activity 6.13 helps students to easily understand the concept of rancidity. This activity can be done by students out of class. Encourage them to report their findings. The following suggested answers help you to harmonize the ideas developed through discussion.

The students should obtain two packets of fried potato chips from the market. They should open one packet and smell it. Half the contents of that packet are placed in a tightly closed container and the other half in an open container. The second packet is kept is. After 15 days, the students should smell the chips kept in the open container, the closed container and the sealed packet. The students should record their observations and find out what is the best way to keep such items.

Soaps and Detergents

You may use the case-study method of teaching for Activity 6.14, which to makes students appreciate the efforts that have been made in the past to invent our modern soap and detergent products. You are required to help the students in forming groups, clarifying the questions in the activity, where required, and facilitating discussion among group members. Let them think independently and share their opinions with their group members. Finally, encourage them to report their common understanding. Match the ideas developed through discussion, based on the following suggested answers.

The students may collect the information regarding the substances which were used for cleaning before the discovery of soaps. The Egyptians used to make a soap-like substance by mixing animal fats with alkaline salts. Incidentally, they used this mixture for treating skin diseases. This is a very fascinating piece of information, as soap is quite responsible for removing horrible disease-causing agents from our skin. The existence of soaps can also be traced from the time of Moses. He specifically gave detailed laws regarding cleanliness and religious purity. The people of his time were also familiar with this soap, like substance, which they used for their cleanliness.

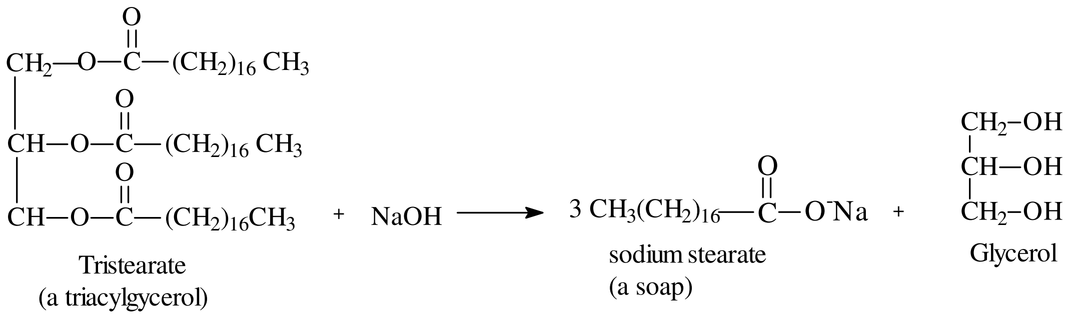
Now let's come to the time of the Greeks. As we all know, they were people of great artistic taste. They used to bath themselves with clay rather than soap, and then they used to anoint themselves with oil, before scraping it with metallic strigil (a skin scraper used in ancient Greece). This used to serve the purpose of soap. The Greeks also used the combination of oil and ashes for washing their clothes. Soap, as we know it today, is more than a cleaning device. Different varieties of soap are now used for purposes like bathing, laundering, shaving and shampooing. Many industries make huge sums by bringing changes to soap making while the actual ingredients are the same as those that were used in ancient times.

Experiment 6.2 is designed to introduce the students to the laboratory method of preparing ordinary soap by the alkaline hydrolysis of fats and oils (saponification).You may arrange a special lab session for this experiment. Refer to the students' textbook for the detailed procedures as well as the apparatus and chemicals required for performing this experiment.

You may apply the predict-observe-explain methodology which is helpful in order to demonstrate this experiment to the students**.** Ask the students to predict the result of this demonstration and give them the time and opportunity to actively think about this question which, in turn, gives them more motivation to study. Then, students can observe the demonstration and contrast the outcome with their prediction(s). Finally, the result is explained. The performance of demonstrations helps the students to learn to think constructively and critically.

The answers to the questions listed under observation and analysis are given below.

1. In this reaction, the animal fat (tallow) or vegetable oil undergoes alkali hydrolysis and coverts to soap.
2. The role of sodium hydroxide is to hydrolyze and convert animal fat (tallow) to soap.
3. The chemical reaction for this reaction.



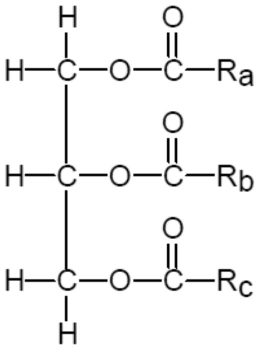
$$\begin{array}{l} CH_2-O-\overset{O}{\overset{\|}{C}}-(CH_2)_{16}\,CH_3 \\ | \\ CH-O-\overset{O}{\overset{\|}{C}}-(CH_2)_{16}\,CH_3 \\ | \\ CH-O-\overset{O}{\overset{\|}{C}}-(CH_2)_{16}CH_3 \end{array} + NaOH \longrightarrow 3\,CH_3(CH_2)_{16}-\overset{O}{\overset{\|}{C}}-O^-Na + \begin{array}{l} CH_2-OH \\ | \\ CH-OH \\ | \\ CH_2-OH \end{array}$$

4. From this experiment, students may learn the following points.
    - Soaps are salts of long-chained fatty esters. They are completely miscible in water.
    - Soaps exist as micelles which are spherical clusters of carboxylate anions that are packed together with the nonpolar hydrocarbon chains in the interior. The sodium ions are scattered throughout the aqueous layer.
    - Because the surfaces of a micelle are negatively charged, individual micelles repel each other and remain dispersed through the aqueous phase.

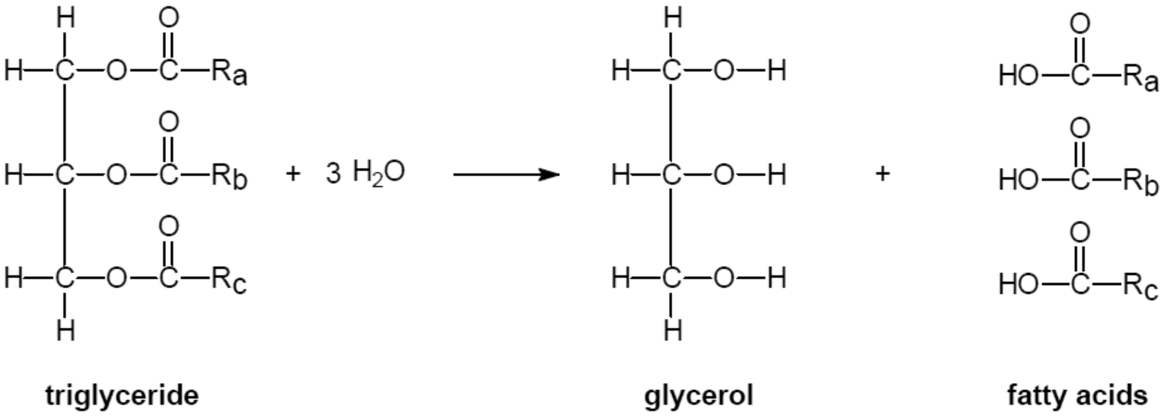
You may create an activity to involve students in discussion. After summing up the unit, you may help students with the difficulties they have in the review exercises.

Additional Notes

The terminology applied to fats is based on the chemical structure of their molecules. Fats and oils belong to a group of biological substances called lipids. Lipids are biological chemicals that do not dissolve in water. They serve a variety of functions in organisms, such as regulatory messengers (hormones), structural components of membranes, and as energy storehouses. Fats and oils generally function in the latter capacity. Fats differ from oils only in that they are solid at room temperature, while oils are liquid. Fats and oils share a common molecular structure, which is represented by the formula below.



This structural formula shows that fats and oils contain three ester functional groups. Fats and oils are esters of the tri-alcohol, glycerol (or glycerine). Therefore, fats and oils are commonly called *triglycerides*, although a more accurate name is triacylglycerols. One of the reactions of triglycerides is hydrolysis of the ester groups.



This hydrolysis reaction produces glycerol and fatty acids, which are carboxylic acids derived from fats and oils. In the fatty acids, R_a, R_b, and R_c represent groups of carbon and hydrogen atoms in which the carbon atoms are attached to each other in an unbranched chain.

The hydrolysis reaction is promoted by acids and by bases. When a strong base such as NaOH (lye) is used, the product contains salts of the fatty acids. These salts of fatty acids are the functional ingredient in soap. The ingredient lists of some soaps include sodium tallowate, a generic name for the mixture of fatty acid salts obtained from tallow (animal fat), and sodium cocoate, for those obtained from coconut oil.

Triglyceride molecules contain mostly carbon and hydrogen atoms, with only six oxygen atoms per molecule. This means that fats and oils are highly reduced (that is, un-oxidized). They are, in this way, similar to the hydrocarbons in petroleum, and like petroleum they are good fuels. The main biological function of triglycerides is as a fuel.

The normal human body stores sufficient energy in fat for several weeks' survival. This storage ability helps the organism deal with unpredictable variations in the food supply. Plants, too, store energy-in fats and oils. Oils are particularly common in seeds, where the stored energy helps seedlings during germination, until they can exploit solar energy through photosynthesis.

Fatty acids contain an even number of carbon atoms, from 4 to 36, bonded in an unbranched chain. Most of the bonds between carbon atoms are single bonds. If all of these bonds are single bonds, the fatty acid is said to be saturated, because the number of atoms attached to each carbon atom is the maximum of four. If some of the bonds between carbon atoms are double bonds, then the fatty acid is unsaturated.

Assessment

You are required to assess students' active involvement in each of the activities. Your student-performance record list is helpful in all your assessments. Don't forget to praise students working above the minimum requirement level. You are required to assist those working below the minimum requirement level. Students working at the minimum requirement level will be able to define fats and oils, write the general structural formula for fats and oils, describe physical properties of fats and oils, and explain the hardening of oils (process of converting oils to hard fats), and explain rancidity.

When you praise students working above the minimum requirement level, recognize their achievements and encourage them to continue working hard and to not become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class.

Additional Questions

Complete the table below by inserting the missing items (a – e).

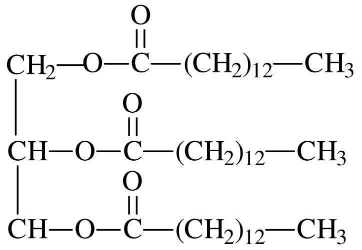
Fatty acid	Elementary Formula	Constitutional Formula
Lauric	$C_{12}H_{24}O_2$	(a)
(b)	$C_{14}H_{28}O_2$	$CH_3(CH_2)_{12}COOH$
Palmitic	*(c)	$CH_3(CH_2)_{14}COOH$
(d)	$C_{18}H_{36}O_2$	$CH_3(CH_2)_{16}COOH$
Oleic	$C_{18}H_{34}O_2$	(e)
Linolic	(f)	$CH_3(CH_2)_{12}(CH)_4COOH$
(g)	$C_{18}H_{30}O_2$	$CH_3(CH_2)_{10}(CH)_6$

Answers to Additional Questions

Fatty acid	Elementary Formula	Constitutional Formula
Lauric	$C_{12}H_{24}O_2$	$CH_3(CH_2)_{10}COOH$
Myristic	$C_{14}H_{28}O_2$	$CH_3(CH_2)_{12}COOH$
Palmitic	$C_{16}H_{32}O_2$	$CH_3(CH_2)_{14}COOH$
Stearic	$C_{18}H_{36}O_2$	$CH_3(CH_2)_{16}COOH$
Oleic	$C_{18}H_{34}O_2$	$CH_3(CH_2)_{14}(CH)_2COOH$
Linolic	$C_{18}H_{32}O_2$	$CH_3(CH_2)_{12}(CH)_4COOH$
Linolenic	$C_{18}H_{30}O_2$	$CH_3(CH_2)_{10}(CH)_6$

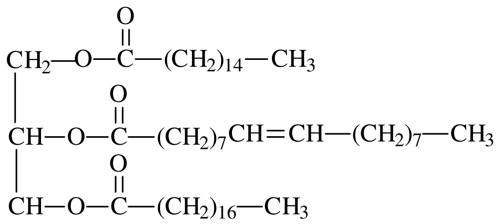
Answers to Exercise 6.7

(a)



Glyceryl trimyristate

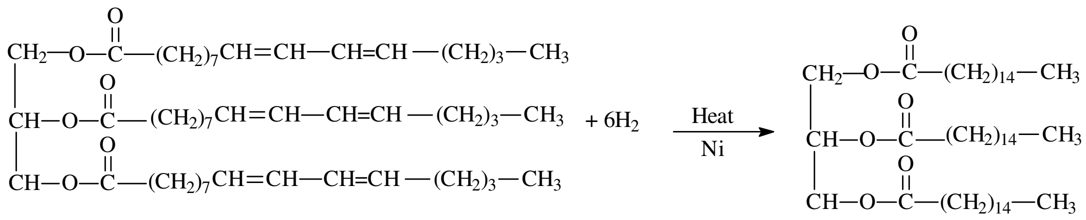
(b)



Glyceryl pamitooleostrate

Answers to Exercise 6.8

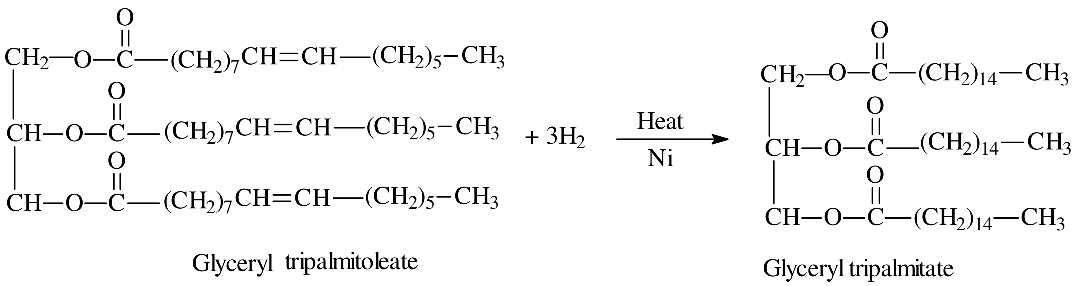
(a)



Glyceryl trilinoleate

Glyceryl tripalmitate

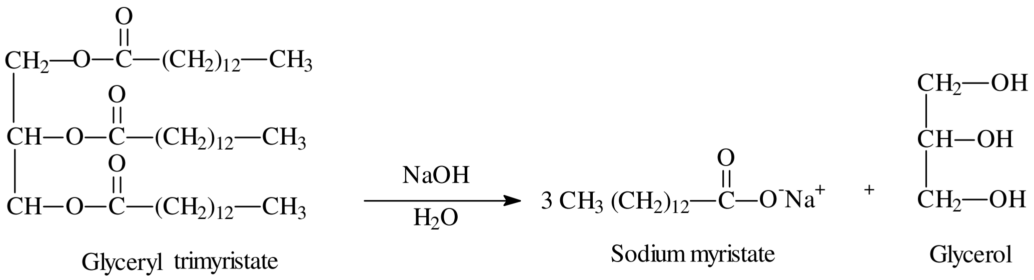
(b)



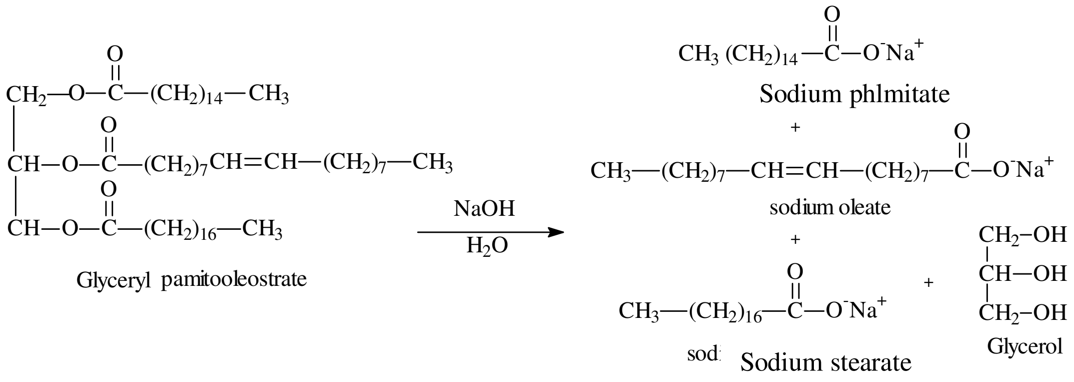
Answers to Exercise 6.9

The base hydrolysis (saponification) products below would be obtained when mixing the following oils with NaOH.

1.



2.



Answers to Review Exercise on Unit 6

Part I: Multiple Choice Questions

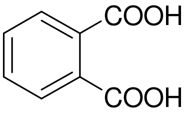
1. B 2. C 3. A 4. B 5. D 6. C 7. A
8. C 9. B 10. C

Part II: Answer the following questions

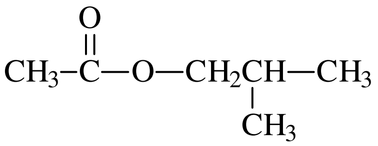
11. The structures for the compounds are:

a.	3-bromobutanoic acid	$CH_3CH(Br)CH_2COOH$
b.	2-hydroxy-2-methylpropanoic acid	$CH_3)_2C(OH)COOH$
c.	2-butynoic acid	$CH_3C\equiv CCOOH$
d.	benzoic acid	$C_6H_5\text{-}COOH$

12. The names of the compounds are:

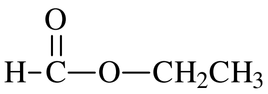
a.	$HCOOH$	Formic acid (Methanoic acid)
b.	$(CH_3)_2CHCOOH$	Isobutanoic acid (2-methyl propanoic acid)
c.	$HOOCCH_2CH_2COOH$	Succinic acid (Butanedioc acid)
d)		Phthalic acid (Benzene-1,2-dicarboxylic acid)
e.	$CH_3(CH_2)_{14}COOH$	Stearic acid (Octadecanoic acid)

13. The structures for the names are:

a) 

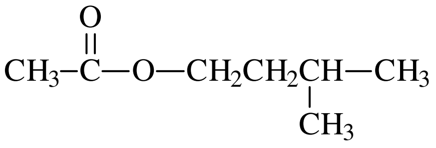
$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2\underset{\underset{CH_3}{|}}{CH}-CH_3$

Isobutyl acetate

b) 

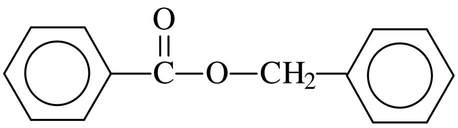
$H-\overset{\overset{O}{\|}}{C}-O-CH_2CH_3$

Ethyl formate

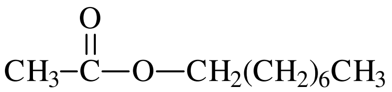
c) 

$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2CH_2\underset{\underset{CH_3}{|}}{CH}-CH_3$

Isopentyl acetate

d) 

benzyl benzoate

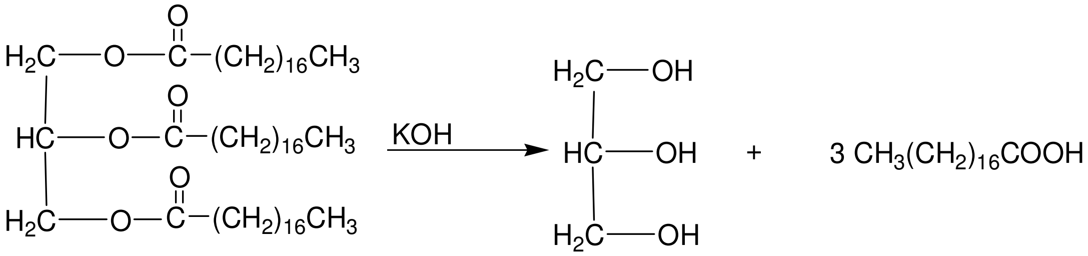
e) 

$CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2(CH_2)_6CH_3$

Octyl ethanoate

14. a. Isopentyl alcohol and acetic acid
   b. Benzyl alcohol and benzoic acid
   c. n-Octanol and ethanoic acid
   d. Isobutyl alcohol and acetic acid
   e. Ethanol and formic acid

15.



16. Define the following and give examples based on the students textbook.

a. **Hydrolysis**

Hydrolysis is a chemical reaction during which molecules of water (H_2O) are split into hydrogen cations (H^+) (conventionally referred to as protons) and hydroxide anions (OH^-) in the process of a chemical mechanism. It is simply the reaction of a substance with water.

b. **Saponification**

Saponification is the hydrolysis of an ester under basic conditions to form an alcohol and the salt of a carboxylic acid (carboxylates). It is commonly used to refer to the reaction of a metallic alkali (base) with a fat or oil to form soap.

c. **Esterification**

Esterification is the general name for a chemical reaction in which two reactants (typically an alcohol and an acid) form an ester as the reaction product. It is a reversible reaction.

d. **Hydrogenation**

Hydrogenation, a form of chemical reduction, is a chemical reaction between molecular hydrogen (H_2) and another compound or element, usually in the presence of a catalyst. The process is commonly employed to reduce or saturate organic compounds. Hydrogenation typically constitutes the addition of pairs of hydrogen atoms to a molecule, generally an alkene. Catalysts are required for the reaction to be usable. Non-catalytic hydrogenation takes place only at very high temperatures. Hydrogen adds to double and triple bonds in hydrocarbons.

e. **Lipids**

Lipids are a broad group of naturally occurring molecules which include fats, waxes, sterols, fat-soluble vitamins (such as vitamins A, D, E and K), monoglycerides, diglycerides, phospholipids, and others. The main biological functions of lipids include energy storage, as structural components of cell membranes, and as important signaling molecules. Although the term *lipid* is sometimes used as a synonym for fats, fats are a subgroup of lipids called triglycerides. Lipids also encompass molecules such as fatty acids and their derivatives

f. Triglycerides

A tiglyceride (triacylglycerol, TAG or triacylglyceride) is an ester composed of a glycerol bound to three fatty acids. It is the main constituent of vegetable oil and animal fats. Most of the fats digested by humans are triglycerides. Triglycerides are formed from a single molecule of glycerol, combined with three molecules of fatty acid. The glycerol molecule has three hydroxyl (OH-) groups. Each fatty acid has a carboxyl group (COOH-). In triglycerides, the hydroxyl groups of the glycerol join the carboxyl groups of the fatty acid to form ester bonds.

g. Rancidity

Rancidity is the chemical decomposition of fats, oils and other lipids. There are three basic types of rancidity. Hydrolytic rancidity occurs when water splits fatty acid chains away from the glycerol backbone in glycerides. Oxidative rancidity occurs when the double bonds of an unsaturated fatty acid react chemically with oxygen. Microbial rancidity refers to a process in which micro-organisms such as bacteria use their enzymes, including lipases, to break down chemical structures in the fat. In each case, these chemical reactions result in undesirable odors and flavors.

h. Hardening of oils

This process converts unsaturated esters of fatty acids into (more solid) saturated esters by hydrogenation using a nickel catalyst. It is used in the manufacture of margarine from vegetable oils.

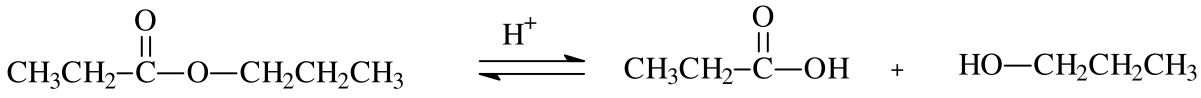
17.

a. Propan-1-ol converted into propanoic acid by oxidation.

The alcohol is heated under reflux with an excess of the oxidizing agent. When the reaction is complete, the carboxylic acid is distilled off.

b. Propyl propanoate into propan-1-ol.

Hydrolysis of propyl propanoate gives propan-1ol and propanoic acid. The reaction is catalysed by dilute acid, and so the ester is heated under reflux with a dilute acid like dilute hydrochloric acid or dilute sulphuric acid.



$$CH_3CH_2-\overset{\overset{O}{||}}{C}-O-CH_2CH_2CH_3 \xrightleftharpoons{H^+} CH_3CH_2-\overset{\overset{O}{||}}{C}-OH + HO-CH_2CH_2CH_3$$

Propyl propanoate Propanoic acid Propan-1-ol

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MINIMUM LEARNING COMPETENCIES (MLCs)

No	Areas of Competencies	MINIMUM LEARNING COMPETENCIES (MLCs)
1	**Fundamental Concepts in Chemistry**	**The students should be able to:** • define chemistry and describe the major fields of chemistry • describe the SI units and their prefixes and write the names and symbols of derived SI units • identify causes of uncertainty in measurement • explain systematic and random errors • define precision and accuracy and analyses a given data in terms of precision and accuracy • define significant figures, determine the number of significant figures in a calculated result and use the scientific notation in writing very large or very small numbers • define scientific method, describe its major steps and use it in solving problems • demonstrate some experimental skills in chemistry and describe the procedures of writing laboratory report
2	**Structure of Substance and Periodic Table**	• narrate historical development of atomic nature of substance • state postulates of Dalton's and modern atomic theories • discuss the discovery of electron & describe the properties of cathode rays • define the terms : radioactivity, radioactive decay and radioisotope • describe the common types of radioactive emission • describe the alpha scattering experiment • describe makeup of the nucleus • explain atomic number and mass number; and define atomic mass and isotope • tell the number of protons, electrons and determine the number of neutrons from the given atomic number and mass number of an element • calculate the relative atomic mass of naturally occurring isotopic elements • characterize electromagnetic radiation in terms of wave length, frequency and calculate the wave length and frequency of the electromagnetic radiation • explain that light has both wave and particle nature and the emission spectra of atoms consist of series of lines • state Bohr's assumption of energy of electron in hydrogen atom

		• calculate the radius of electron orbit, the electron velocity and the energy of an electron using Bohr model • explain that atoms emit or absorb energy when they undergo transition from one stats to another • explain that the spectrum of hydrogen demonstrates the quantized nature of the energy of its electron • explain the short coming of Bohr's theory • state Hinesburg's uncertainty principle and describe the significance of electron probability distribution • explain the quantum numbers n_l, m_l and m_s • write all possible sets of quantum numbers of electrons in an atom • describe the shapes of orbital designated by s, p, and d • explain Aufbau principle Pauli exclusion principle and Hundi's rule • write ground state electron configuration of elements with the periodicity of elements • classify elements as representative, transition and inner-transition • explain the general trends in atomic radius, ionization energy, electron affinity, electro-negativity and metallic character of elements within a period and a group of the periodic table • write the advantages of periodic classification of elements • describe the reason why atom from chemical bonds • define chemical bonding and describe the types of chemical bonding and the mechanisms of bonding process • define ionic bonding and explain the formation of ionic bonding • give examples of ionic compounds and describe their properties • define lattice energy • calculate the lattice energy of ionic crystals from a given data using the methods of Born-Haber cycle • discuss the limitation of Octet rule • describe ionic bonding using Lewis electron dot symbols • define Covalent bonding and explain the formation of Covalent bonding • give examples of covalent molecules & describe their properties • draw Lewis structures of some covalent molecules • define resonance and draw resonance structures of some covalent molecules and polyatomic ions • discuss the exceptions to the Octet rule in covalent bonding • illustrate the formation of coordinate covalent bonding using examples

		• distinguish between polar and non-polar covalent molecules • describe the properties of covalent molecules • define intermolecular force and name the different types of intermolecular forces • explain dipole-dipole interaction and give examples of them • define hydrogen bonding and explain the effects of it on the properties of substances • explain dispersion forces and give examples of them • define metallic bonding and explain the properties of metals related to the concept of bonding • name two chemical bond theories • explain the VBT and distinguish between the Lewis model and valence bond model • discuss the overlapping of orbital in covalent bond formation; • explain hybridization and show its process involved in some covalent molecules • draw hybridization diagram for the formation of $sp,sp^2,sp^3,sp^3d,sp^3d^2$ hybrids • discuss the hybridization involved in compounds containing multiple bonds • explain the MOT and describe molecular orbital using atomic orbital and bonding and anti-bonding orbital • write the electron configuration of simple molecules using molecular orbital model • define bond order of some simple molecules and their stability using bond order • describe the Valence Shell Electron Pair Repulsion(VSEPR) theory • define the bonding pairs ad non-bonding pairs of electron • describe how electron pair arrangement of molecules can be predicted from the number of electron pairs • explain the term dipole moment with the help of a diagram • predict the geometrical shapes of some simple molecules on the basis of hybridization and the nature of electron pairs • explain that the VSEPR theory is used to explain the molecular parameters like bond angle and polarity of the molecule • construct models to represent shapes of some simple molecules • define crystal, name four types of crystalline solids, give examples and describe their properties • mention the types of attractive forces that exist within each type of crystalline solids

3	**Chemical Reaction**	• define chemical kinetics and reaction rate • perform an activity to measure rate of reaction • solve problems related to the rate of reaction • list five factors that affect the reaction rate and explain how they affected by giving examples • define catalyst, positive catalyst and negative catalyst • distinguish between homogeneous and heterogeneous catalysts, biological and non-biological catalysts and homogeneous and heterogeneous reactions • perform an activity to show the effect of nature of reactant on surface area, concentration, temperature and catalyst on the reaction rate • state collision theory and transition state theory, and describe how these theory can be used to explain change in reaction rate • define activation energy, activated complex, rate law, order of reaction rate construct and half-life • sketch and label the energy profiles of reactions which are exothermic and endothermic • determine reaction order and calculate rate constants from a given experimental data • explain the zero, 1st and 2nd order reaction using concentration verses time curve • calculate the half-lives of zero, 1st, and 2nd orders from experimental data • explain reaction mechanism, molecularity of a reaction, and rate determining stop and give specific examples to illustrate rate determining step • explain the relationship between the reaction pathway and the rate law and use rate law to suggest possible reaction mechanism for a reaction • explain reversible and irreversible reactions • define dynamic equilibrium, discuss how chemical equilibrium is established and explain its characteristics • state the law of mass action • define equilibrium constant, write its expression involving concentration for chemical reactions and calculate it • write the equilibrium constant expression involving partial pressure for chemical reactions and calculate it • show relationship between K_c and K_p • distinguish between homogeneous and heterogeneous equilibrium reaction • define reaction quotient and use equilibrium quotient to predict the

		direction of the reaction and position of equilibrium • calculate the equilibrium concentrations of a given initial concentrations • list factors that affect chemical equilibrium • state Le-chatlier principle and use it to explain the effect of changes in temperature, pressure, concentration and presence of catalyst of a reaction
4	**Carboxylic acids, Esters, Fats, Oils and Polymers**	• write the general formula of saturated monocarboxylic acid • write the molecular formulas and names of the first six members of saturated monocarboxylic acids & give their structural formulas • give examples for monocarboxylic, dicarboxylic and tricarboxylic acids and name some branched carboxylic acids • explain the general methods of preparation of saturated monocarboxylic acids and the industrial and laboratory preparation of acetic acid • conduct an experiment to prepare acetic acid in the laboratory; • explain why “Tella” or “Tej” turns sour • describe physical properties and chemical reactions of saturated monocarboxylic acids • name and write structural formulas of some fatty acids • list common sources of esters • write the general formulas of esters and the molecular formulas and names of the first six members of esters • explain the method of preparation of esters • describe physical properties and chemical reactions of esters • define fats and oil • write the general formula of fats and oil and the structures for some common triglycerides • describe physical properties of fats and oil • explain rancidity and hardening of oil • define soap and detergent • explain saponification and the cleaning action of soap • prepare soap

Federal Democratic Republic of Ethiopia
Ministry of Education

Chemistry Syllabus, Grades 11

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 stards.

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 **“stard level”**. Students working at the stard 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 fore 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

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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 11

Unit	Sub-unit	Number of Periods	
		Sub-unit	Total
Unit 1: Fundamental Concepts in Chemistry	1.1 The scope of chemistry 1.2 Measurements units in chemistry 1.3 Chemistry as experimental science	2 8 5	15
Unit 2: Atomic Structure Periodic Table	2.1 Historical development of the atomic nature of substances 2.2 Dalton's atomic theory the modern atomic theory 2.3 Early experiments to characteristics the atom 2.4 Makeup of the nucleus 2.5 Electromagnetic radiation (EMR) atomic spectra 2.6 The quantum mechanical model of the atom 2.7 Electronic configurations orbital diagrams 2.8 Electronic configurations the periodic table of the elements	1 2 3 2 9 5 2 4	28
Unit 3: Chemical Bonding Structure	3.1 Introduction 3.2 Ionic bonding 3.3 Covalent bonding 3.4 Metallic bonding 3.5 Chemical bonding 3.6 Types of crystal	1 5 15 2 8 1	32
Unit 4: Chemical Kinetics	4.1 Rate of reaction 4.2 Theories of reaction rates 4.3 Rate equation or rate 4.4 Reaction mechanism	5 4 7 3	19
Unit 5: Chemical Equilibrium Phase Equilibrium	5.1 Chemical equilibrium 5.2 Phase equilibrium	14 4	18
Unit 6: Carboxylic acids, Esters, Fats Oils	6.1 Carboxylic acids 6.2 Esters 6.3 Fats oils	8 6 6	20

General Objectives of Grade 11 Chemistry

To develop understanding acquire knowledge of:

- the scope major fields of chemistry;
- the appropriate use of SI units uncertainties in measurement;
- atomic theory structure of the atom;
- the particle wave nature of electrons;
- the relationship between electronic arrangement of the atoms periodic trends of the elements;
- formation of chemical bonding properties of compounds formed by different types of bonds;
- predicting the shapes of molecules intermolecular forces in covalent compounds;
- rates of chemical reactions ways of enhancing controlling it;
- chemical equilibrium in a reversible system methods of shifting it to the desired direction;
- sources, properties, compositions, structures, names uses of organic substances such as organic acids, esters, fats oils.

To develop skills abilities of:

- using scientific methods in solving problems;
- expressing calculated values using appropriate number of decimal places or significant figures scientific notation;
- conducting experiments following laboratory procedures;
- solving problems related to atomic structure chemical equilibrium;
- applying bonding theories to predict properties of substances;
- conducting experiments to observe analyze the physical properties of substance determine the type of bonding;
- determining rates of chemical reactions, amounts of reactants products at any time of a reaction;
- demonstrating experiments to investigate the behaviour of different equilibrium systems to explain effects of some factors on equilibrium system;
- carrying out laboratory activities involving carboxylic acids, esters, fats oils;
- preparing soap testing its cleaning action in different types of water;
- design conduct simple experiments appropriate to their level.

To develop habits attitudes of:

- appreciating the importance of chemical kinetics in industrial production living organisms;
- practicing scientific methods of solving problems;
- appreciating the interrelationship between electronic structure of the atoms periodic classification of the elements;
- developing personality characteristics such as neatness, exactness, diligence, responsibility carefulness;
- relating a scientific knowledge to everyday application

Unit 1: Fundamental concepts in chemistry *(15 periods)*

Unit outcomes: Students will be able to:

- understand the scope of chemistry;
- select and use appropriate SI units;
- understand the causes of uncertainty in measurement;
- express the result of any calculation involving experimental data to the appropriate number of decimal places or significant figures;
- use scientific methods in solving problems;
- demonstrate an understanding of experimental skills in chemistry;
- demonstrate a knowledge of laboratory procedures;
- describe scientific enquiry skills along this unit: observing, inferring, predicting, comparing and contrasting, communicating, analyzing, classifying, applying, theorizing, measuring, asking question, developing hypothesis, designing experiment, interpreting data, drawing conclusion, making generalizations problem solving.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**1. Fundamental concepts in chemistry** **1.1 The scope of chemistry** (2 periods)	
• Define chemistry	• Definition of chemistry	Students should be able to define chemistry as the study of the properties of substances. Students could be asked to exp on this definition by suggesting what they interpret this definition to mean. These could include: Some of the physical properties of substances The stability of substances The composition of substances The synthesis of substances The chemical reactions of substances The uses of substances
• Distinguish the major fields of chemistry	• Major fields of chemistry	Students should appreciate that chemistry is divided into four major branches be able to give a brief account of each. Inorganic chemistry – the chemistry of the elements their compounds including carbon monoxide, carbon dioxide, carbonates hydrogen carbonates but excluding all other compounds of carbon. Organic chemistry – the chemistry of carbon compounds except for carbon monoxide, carbon

Competencies	*Contents*	*Suggested Activities*
		dioxide, carbonates hydrogen carbonates. Physical chemistry – the chemistry of the structure physical properties of substances how these are related, the factors that affect chemical reactions. Analytical chemistry – the chemistry of qualitative identification quantitative determination of the constituents of matter There is also another branch of chemistry which bridges chemistry biology. Biochemistry is the chemistry of living organisms. Students could investigate how each of these major branches of chemistry can be further subdivided e.g.
• Distinguish the subdivision of branches of chemistry		• Inorganic chemistry – Coordination chemistry, etc. • Organic chemistry – Synthetic organic chemistry, etc. • Physical chemistry – Chemical kinetics, thermodynamics, etc. • Analytical chemistry – Qualitative analysis, quantitative analysis, etc.
	1.2 Measurements units in Chemistry (8 periods)	
• List describe the seven SI units their prefixes	• SI units (The International system of units) - Basic SI units	Students should be aware that a common group of units called S.I. units are used by scientists, including chemists, throughout the world. Students could discuss why having a common group of units is an advantage e.g. what is the problem with measuring pressure in Pascal, bars, pounds per square inch millimetres of mercury? Students should understand that there are seven base S.I. units. Those relevant to chemistry are: Length – metre – m

Competencies	*Contents*	*Suggested Activities*
		Mass – kilogram – kg Time – second – s Electric current – ampere – A Temperature – Kelvin – K Amount of substance – mole – mol
• Write the names symbols of derived SI units	- Derived SI units • Common prefixes used in SI units	Students should understand that other units are derived from them. Those relevant to chemistry include: kg/m^3, g/cm^3 (density) mol/dm^3(concentration) N/m^2 = Pa (pressure) Students should be aware that temperature is usually measured in degrees Celsius or centigrade, °C, although this is not the S.I. unit. Students should appreciate that it is impractical to use the same unit for expressing all values for a particular quantity. Students could discuss why it would be difficult to express the amounts of chemicals used in industrial quantities the amounts of chemicals present in pharmaceutical preparations both in kg. From this discussion the teacher should develop the idea of using prefixes to express multiples of units: • mega – M - 10^6 • kilo – k - 10^3 • milli – m – 10^{-3} • micro - μ - 10^{-6}
• Use the factor label method for solving problems making conversion of SI units.	\	Students could practice conversion between units e.g. between micrograms, milligrams, grams, kilograms, mega grams (tonnes). Students should be made aware of the appropriate use of upper lower case e.g. possible confusion between mega milli.
• Describe uncertainty of measurement • Identify the digits that are certain the ones that are uncertain given a number representing	• Uncertainty in measurements	Students should appreciate that, even with careful measurement, there is a limit to the degree of accuracy of an instrument. For example, with skill practice a student may read a burette to the nearest 0.05 cm^3 but no more

Competencies	*Contents*	*Suggested Activities*
measurement • Identify causes of uncertainty in measurement • Define precision accuracy • Estimate the precision possible for any instrument they use in the laboratory • Explain system rom errors • Analyse given data in terms of precision accuracy	• Precision accuracy	accurately than that. Students should have realistic expectations of the accuracy of the instruments they use. Students should have an awareness of the potential sources of errors in taking readings e.g. measuring the temperature of a solution by inserting a thermometer but waiting sufficient time for the thermometer to attain the same temperature as the solution. Students should be able to identify an anomalous reading in a set of data suggest a reason for it. Students should appreciate the limitations of accuracy in taking measurements. For example, if the mass of a chemical is given as 2.5 g the actual mass may be anything from 2.450 g to 2.549 g. Students should be able to express the uncertainty in any measurement the limit of accuracy of a calculation using this measurement.
Define decimal places Determine the number of decimal places in a calculated result	• Decimal places	Students should understand that numbers are rounded up either to a particular number of decimal places or significant figures. The process for this is similar. Students should know that: to express a value to the nth decimal place they look at the value of the n+1th digit if the n+1th digit is 5 or more they round up the nth digit if the n+1th digit is 4 or less they leave the nth digit unchanged Students should be talked through some examples then attempt some examples for themselves e.g. express 24.546 g to 1 decimal place

Competencies	*Contents*	*Suggested Activities*
Define significant figures Determine the number of significant figures in a calculated result	• Significant figures	the second decimal place digit is 4 so do not round up expressed to 1 decimal place the number is 24.5 g Similarly: express 0.167 mol/dm^3 to 2 decimal places the third decimal place digit is 7 so round up expressed to 2 decimal places the number is 0.17 mol/dm^3 Students should know that a similar system is used for significant figures: to express a value to the nth significant figure they look at the value of the n+1th digit if the n+1th digit is 5 or more they round up the nth digit if the n+1th digit is 4 or less they leave the nth digit unchanged Students should be talked through some examples then attempt some examples for themselves e.g. express 348 °C to 2 significant figures the third significant digit is 8 so round up expressed to 2 significant figures the number is 350 °C Similarly: express 4.381 g to 3 significant figures the fourth significant digit is 1 so do not round up expressed to 3 significant figures the number is 4.38 g Students should understand that it is not possible to express an answer to a higher degree of accuracy that the data from which it is derived. For example: The mass of a metal block = 15.2 g The volume of the metal block = 5.6 cm^3 The density of the metal = mass/volume = 2.714857 But the mass the volume are expressed only to one decimal

Competencies	*Contents*	*Suggested Activities*
		point so the density cannot be given any more accurately than one decimal point = 2.7 g/cm^3
Use the scientific notation in writing very large or very small numbers	- Scientific notation	Students should appreciate the problems of writing very large or very small numbers in terms of the numbers of zeros involved e.g. 2300000, 0.000000154. They should understand that scientific notation involves writing numbers to a common format in which there is: One digit to the left of the decimal point The remaining digits to the right of the decimal point The number is multiplied by a power of ten The power of ten is determined by number of places the decimal point must be moved to the left or right regenerate the number in its original format For example: 2300000 = 2.3 x 10^6 0.000000154 = 1.54 x 10^{-7} Students should practice writing very large very small numbers in scientific format.
	1.3 Chemistry as experimental Science (5 periods)	
Define scientific method Describe the major steps of the scientific method Use scientific methods in solving problems	The scientific method	Students should appreciate that there are certain important steps to applying the scientific method. Identifying the problem - gathering information, devising a solution Carrying out experiments – observation, measurement, recording, inferring Examining the results of the experiment – analysis, generalisations, synthesis Using conclusions – explanation, prediction Further experiments – checking predictions, accounting for

Competencies	***Contents***	***Suggested Activities***
		unexpected results
Demonstrate some experimental skills in chemistry	Some experimental skills in chemistry	Students should learn the skills needed to carry out experimental work successfully. These should include: • Manipulation of apparatus • Accurate measurement • Careful observation • Safe practice with regard to themselves others Students should learn to use common apparatus with care safety. This should include: Reaction vessels such as beakers, flasks, boiling tubes test tubes Measuring equipments such as pipettes, burettes, balances thermometers Support heating devices such as st clamp, tripod gauze, spirit burner Bunsen burner
Describe the procedures of writing laboratory report	Writing a laboratory report	Students should learn how to write an accurate comprehensive report on each piece of work carried out in the laboratory. This should include: • An explanation of the method used • The apparatus used including diagrams if necessary • Accurate observations • Use of tables to record data • Conclusions based on observations • Use of graphs or charts to present data • Identification of any patterns which may be evident • Inclusion of chemical equations where appropriate • Calculations if appropriate Students could be given some practical activities to carry out where emphasis is placed on scientific method.

Assessment

The teacher should assess each student's work continuously over the whole unit 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 chemistry and distinguish the major fields of chemistry the subdivision of branches of chemistry; describe the SI units and their prefixes; write the names and symbols of derived SI units; identify causes of uncertainty in measurement; define precision and accuracy and analyze given data in terms of them; explain system and rom errors; define decimal places and determine the number of decimal places in a calculated result; define significant figures and determine the number of significant figures in a calculated result; use the scientific notation in writing very small or very large numbers; define scientific method and, describe the major steps of the scientific method and use scientific methods in solving problems; demonstrate some experimental skills in chemistry and describe the producers of writing laboratory report.

Students above minimum requirement level

Students working above the minimum requirement level should be praised their achievements recognized. They should be encouraged to continue working hard 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 additional lesson time during breaks or at the end of the day.

Unit 2: Atomic Structure and Periodic Table *(28 Periods)*

Unit outcomes: students will be able to:

- understand the historical development of atomic structure;
- explain the experimental observations inferences made by some famous scientists to characterize the atom;
- list describe the subatomic particles;
- explain the terms atomic mass isotope;
- understand electromagnetic radiation, atomic spectra Bohr models of the atom;
- do calculations involving atomic structure;
- describe the quantum mechanical model of the atom the related postulates principles;
- demonstrate an understand of periodic law how electronic configurations of atoms are related to the orbital diagrams can explain periodic trends;
- describe scientific enquiry skills along this unit: inferring, predicting, classifying, comparing contrasting, communicating, asking questions making generalization.

Competencies	*Contents*	*Suggested Activities*
Students will be able to:	**2. Atomic Structure Periodic Table**	
• State briefly the history of development of atomic nature of substances	**2.1 Historical development of the atomic nature of substances *(1 period)***	Students should be able to appreciate: early Greek concepts of the composition of matter based on visionary beliefs but not on evidence (Democritus Vs Zeno of Elea)
• State postulates of Dalton's Atomic Theory	**2.2 Dalton's Atomic Theory the Modern Atomic Theory** *(2 periods)* • Postulates of Dalton's Atomic Theory	why Dalton's Atomic theory is considered scientific even though it had errors how scientific ideas develop based on previous scientific findings (in this case the law of conservation of mass by Lavoisier, the law of definite composition by Proust the law of multiple proportions by Dalton himself the first modern list of elements published by Lavoisier)
	• Postulates of Modern Atomic Theory	Students need to see the Postulates of Dalton's Atomic theory as they were formulated by Dalton himself then the modified version of it(the so called Modern Atomic Theory) as it stands today
• State Postulates of Modern Atomic Theory	• How the theory explains the mass laws	The mass laws (law of conservation of mass, law of definite composition, law of multiple proportions) should be discussed together so that students

Competencies	*Contents*	*Suggested Activities*
• State the laws of conservation of mass, definite proportions multiple proportions the basis of each		can see the connection between each one of them. You may act as the Devil's advocate by telling the students that you don't believe in the law of conservation of mass by raising issues the students are most likely to raise as ' evidences' against the law. The mass of a piece of wood before after it is burnt, in an ordinary stove, to ashes is not the same- showing that mass can be created. A seedling growing shows increase in mass which can be taken as 'evidence' that mass can increase during a biochemical reaction. Then work backwards to show them that indeed mass is conserved by showing them if one burned the wooden log inside a completely closed chamber, that does not allow any gases or otherwise to leave the chamber, then there would not be mass loss. Look for more similar examples that the students take as 'evidences' against the law guide them so that they correct the misconceptions they earlier formulated, on their own. Guide the students understand how Dalton's postulates explain the mass laws one by one.
• Discuss the discovery of electron • Describe the properties of cathode rays • Define the terms: radioactivity, radioactive decay radio-isotope. • Describe the common types of radioactive emissions. • Discuss the alpha scattering experiment. • Describe the major contribution of experiments by	**2.3 Early experiments to characterize the atom** (3 periods) • Discovery of the electron • Radioactivity discovery of nucleus • Discovery of neutron	Students should appreciate the experiments carried out how the lessons learned from their findings (the discoveries of fundamental subatomic particles the atomic nucleus) helped in understanding the composition structure of the atom • J.J. Thompson-electron • Ernest Rutherford-nucleus and proton • Chadwick-neutron Students should able to give a brief summary of the famous scientist carried out by Thomson. That is if

Competencies	*Contents*	*Suggested Activities*
Thomson, Millikan Rutherford concerning atomic structure.		two wires are subjected to a high electric potential then brought close together, a spark or arc jumps from one wire to the other. If the two ends sealed in a glass tube (Cathode Ray) that is then highly evacuated, the discharge from one wire to the other is gentler. This discharges called a cathode ray, causes the glass tube to emit a faint yellowish green glow. Students should also appreciate the properties of the cathode ray realize that cathode rays (Electrons) are fundamental particles found in all matter. Students should be aware that though Rutherford, from his alpha scattering experiment, concluded that an atom must possess a very tiny nucleus that contains all of the positive charge in the atom nearly all its mass, this is the only way he could account for the scattering of some α-particles at large angles from thin metal foils, the fact that most of the α-particles passed through the foil nearly unaffected. Students should understand that the mass of the nucleus could not be accounted for with only protons this became evident when Chadwick, another student of Rutherford, discovered a nuclear particle, without electrical charge-hence named neutron. Note the continuity of ideas (the importance of being in the right place at the right time) considering the pivotal role Rutherford his students played in discovering the constituents of the nucleus observe how the discovery of radioactivity by Becquerel was necessary for the discovery of the nucleus of the atom see how the discovery of the neutron helped discover the presence of isotopes
• Describe makeup of the nucleus	**2.4 Makeup of the nucleus** (2 periods)	Students should be very comfortable in using atomic number, mass number

Competencies	*Contents*	*Suggested Activities*
	- Constituents of the nucleus	atomic mass. Students should know that atomic number was discovered (not protons) by a young British physicist, Mosley in 1913 by correlating x-ray spectrum of each metal with its order in the periodic table (its atomic number) realized that the nuclear charge increased by one unit for each element. In other words atomic number is an element's position number in the periodic table is numerically equal to the number of protons in the nucleus of an atom. Students should know that the total number of nucleons (protons neutrons) found in the nucleus is called mass number.
- Define atomic mass. - Define isotope. - Calculate the relative atomic mass (atomic mass) of naturally occurring isotopic elements.	- Atomic mass isotopes	Numerically mass number atomic mass are very close but they should not be taken to be one the same. Atomic mass is the average of the naturally occurring isotopes of an element weighted according to their abundances. Students need to be given group assignments to calculate the atomic mass of few elements like chlorine, bromine, carbon hydrogen from the masses abundances of the corresponding isotopes. Ask students to find the number of naturally occurring isotopes. Students should be aware that isotopes are different atoms (not identical) of the same element. Analogies would be helpful here to make the concept clear. The word isotope comes from the Mendel eve's periodic law (atomic properties of elements are periodic functions of their atomic masses.) When he tried to arrange the hydrogen atoms that differ in number of protons, hence mass number (or carbon atoms with different masses for that matter) in to different boxes of the Periodic table, he saw that their properties warranted them same position despite what the

Competencies	*Contents*	*Suggested Activities*
		law predicted. They are called isotopes because different atoms (of the same element) occupy the same position in the table. Students should be able to see that isotopes of an element are different atoms not identical. Students should be able to see that our models (mental perceptions) of the atom become more accurate with new findings emerging that help scientists formulate new models that agree with facts waiting to be explained. Students should be able to see that the nature of the nucleus has been known for quite some time since the days of Rutherford his associates. What has been elusive is the position velocity of electrons.
• Characterize Electromagnetic Radiation (EMR) in terms of wave length, frequency speed. • Calculate the wavelength frequency of the EMR.	**2.5 Electromagnetic Radiation (EMR) Atomic Spectra** (9 periods) • Electromagnetic Radiation	Students should have appreciation for electromagnetic radiations express the energies of waves (photons) in terms of wave length (λ) and frequency (ν). Students should carry out calculations on the wavelength frequency of EMR.
• Explain that light has both wave particle nature. • Define photon as a unit of light energy. • Explain how photon theory explains the photoelectric effect the relation between photon absorbed electron released.	• The quantum theory photon	Students should be aware that light has both wave particle nature. This section has to be treated lightly as the concept of waves treated as particles vice versa is difficult for the students. Students should focus on the meaning of the word quantum which is a packet of energy equal to hν. It is the smallest quantity of energy that can be emitted or absorbed. Energy related to emissions or absorptions are whole number multiples of this quantum, $E = nh\nu$.
• Explain that the	• Atomic Spectra	Students should understand that atomic or line spectra (plural of spectrum) are produced from emission but not absorption of

Competencies	*Contents*	*Suggested Activities*
emission spectra of atoms consist of series of lines		distinct frequencies of photons. This implies that the dissociated molecules into atoms have to be excited either thermally or electrically. As excited states are not stable indefinitely they emit radiations (photons) of definite frequencies upon returning to the ground state. A line spectrum that consists of only relatively few wavelengths that is produced from light emitted by excited atoms is a unique feature of an element that can be used for identification purposes.
• State Bohr's assumption of energy of electron in hydrogen atom.	• The Bohr Model of the Hydrogen atom	Students should know that the Rutherford model of an atom could not explain why the emission spectrum of an atom was not a continuous one instead of being discrete. Students should not try to see how the empirical data generated using the Rydberg's formula correlate with Bohr's atomic model for hydrogen.
• Explain that the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron. • Explain that atoms emit or absorb energy when they undergo transitions from one state to another • Explain the short comings of Bohr's theory. • Calculate the radius of electron orbit, the electron velocity the energy of an electron using Bohr model.	• Limitations of the Bohr Theory	Soon after the nuclear model of the atom was proposed, which failed to account for the atomic spectra of elements, Bohr suggested a model for the H-atom that predicted the existence of line spectra. The Bohr model of an atom was successful in accounting for the spectral lines of H-atom, which indicated he was on the right track. His model, however, failed to predict the wavelengths of spectral lines of atoms more complicated than hydrogen demonstrated that there was a basic flaw somewhere in the theory. Students should know the relationship between atomic (line) spectra nd Bohr model of the hydrogen atom. Students should be helped to see what the Bohr model of the atom looks like including what make it a better model than the previous ones its limitations in explaining certain facts. Ask students to draw pictures of elements individually whose atomic number is less than 30 then discuss

Competencies	***Contents***	***Suggested Activities***
		them in groups. Students should carryout calculations on the radius of electron orbit, the electron velocity the energy of an electron using Bohr model.
• State Heinsberg's uncertainty principle • Describe the significance of electron probability distribution	**2.6 The Quantum Mechanical Model of the Atom** (5 periods) • The Heinsberg's Principle	Students should know the breakdown of classical mechanics in relation to subatomic particles like the electrons, hence necessity for a new theory-the Quantum mechanical model of the atom.
• Explain the quantum numbers n, l, m_l, m_s • Write all possible sets of quantum numbers of electrons in an atom.	• Quantum numbers.	Students should understand that according to the quantum mechanical model of the atom an electron is described in terms of 4 quantum numbers, of which three describe the orbital- the space the electron is supposed to occupy the 4th relates to the spin of the electron. Students should appreciate the set of the quantum numbers as a street address of a person in a country. The four quantum numbers are principal quantum number, subsidiary quantum number, magnetic quantum number, spin quantum numbers.
• Describe the shapes of orbital designated by s, p d	• Shapes of atomic orbitals	Students should be aware that there are different shapes of atomic orbital. For example, S-spherical, P-dumbbell. They also differ in size depending upon their energy. It is very important to make students aware of the way scientific ideas progress. At this point ask the students to form groups discuss to see how our beliefs about structure of matter evolved. What was once believed to be continuous is now without doubt particulate (atom). Dalton's atom was considered indivisible; now we know there are fundamental particles inside the atom. It is now clear what is found in the nucleus but there is still some ambiguity about the exact location energy of the electrons that are

Competencies	*Contents*	*Suggested Activities*
		known to move around the nucleus. The era of modern atomic theory begins with Dalton's model of the atom. This should not make the students think that his model is the best of all similar models. Ask students to list down what has been retained in the modern atomic theory from Dalton's postulates. Help students organize a debate to award a prize to one prominent scientist from among the many who contributed to the current model of the structure of the atom. This will give them a chance to revise the unit. You could tell them that you are Manchester United fan in the Engl's football premier league because Dalton was from Manchester area.
• Explain Aufbau principle. • Explain Pauli-exclusion principle. • Explain Hund's rule • Write ground state electron configurations of multi-electron atoms.	**2.7 Electronic configurations orbital diagrams** (2 periods) • Aufbau principle • Pauli-exclusion principle • Hund's rule • Ground state electronic configuration of the elements.	Students need to have appreciation for the following principles to be used in writing electron configurations of atoms along with the quantum numbers • Aufbau's principle • Pauli's Exclusion principle • Hund's principle Giving clear concise definitions for new terms from the beginning simplifies the effort students put to learn complex concepts. You may tell them that these are German words- Aufbau means building order both Pauli Hund are names of scientists. Ground state electronic configurations of the elements can be solved if students are given the atomic number in conjunction with the charge of the atom using Aufbau principle as there is only one of its kinds. However, to check that they know the difference between excited ground states ask students individually to write excited electronic configurations for some elements or ions.

Competencies	*Contents*	*Suggested Activities*
- Correlate the electron configuration of elements with the periodicity of elements. - Give a reasonable explanation for shape of the periodic table. - Classify elements as representative, transition inner-transition elements.	**2.8 Electronic Configurations The Periodic table of The Elements** (4 periods) - The modern periodic table	Do not take it for granted that students know the difference between the Mendeleve's Modern periodic laws. Ask them to state these laws prior to teaching this section. They should be able to use the electronic configurations of elements to locate the position in the periodic table vice versa their properties. Give assignments individually then ask them to iron out the differences, if there are any.
- Explain the general trends in atomic radius, ionization energy, electron affinity, electro negativity metallic character of elements within a period a group of the periodic table.	- Classification of the elements - Periodic properties	Students should be aware that elements can be classified in the periodic table according to their electronic configurations. Students should appreciate the general trends in periodic properties of elements across the period along the group in the periodic table.
- Write the advantages of periodic classification of elements.	- Advantages of periodic classification of elements.	Ask students to study all the chemistry there is for all elements. You can also ask them if they think there is a person who tries to study by heart all the facts summarized in the Periodic table. Can they suggest any tool or means of simplifying this task? If they know the property of an element at the top of a group they can give an educated guess of the property of an element down the group if they understand the trends in change. Students should use the periodic table for further study of elements their compounds.

Assessment

The teacher should assess each student's work continuously over the whole unit 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: narrate the historical development of atomic nature of substances; state postulates of Dalton's atomic theory and modern atomic theory; discuss the discovery of electron and the alpha scattering experiment, and describe the properties of cathode rays; define radioactivity, radioactive decay and radioisotope and describe the common types of radioactive emissions; describe makeup of nucleus; define atomic mass and isotope; calculate the relative atomic mass of naturally occurring isotopic elements; characterize EMR in terms of wavelength, frequency and speed and calculate the wavelength and frequency of the EMR; explain that light has both wave and particle nature; explain that emission spectra of atoms consist of series of lines; state Bohr's assumption of energy of electron in hydrogen atom calculate the radius of electron orbit, the electron velocity and the energy of an electron using Bohr model; explain that the line spectrum of hydrogen demonstrates that quantized nature of the energy of its electron; explain that atoms emit or absorb energy when they undergo transitions from one state to another; explain the short coming of Bohr's theory; state Heinsberg's uncertainty principle and describe the significance of electron probability distribution; explain the quantum numbers n,l,m_l and m_s and write all possible of electrons in an atom; describe the shapes of orbital designated by s,p,d, f ; explain Aufbau principle, Pauli-exclusion principle Hund's rule; write ground state electronic configurations of multielectron atoms correlate the electron configuration of elements with the periodicity of elements; classify elements as representative , transition inner-transition elements; explain the general trends in atomic radius , ionization energy, electron affinity, electronegativity, metallic character of elements within a period a group of the periodic table write the advantages of periodic classification of elements.

Students above minimum requirement level

Students working above the minimum requirement level should be praised their achievements recognized. They should be encouraged to continue working hard 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 additional lesson time during breaks or at the end of the day.

Unit 3: Chemical Bonding and Structure *(32 Periods)*

Unit outcomes: Students will be able to:

- understand that a chemical bond is an attractive force between particles;
- demonstrate an understanding of the formation general properties of substances containing ionic, covalent metallic bond;
- draw Lewis structure for simple ionic covalent compounds;
- understand the origin of polarity within molecules;
- describe the formation nature of hydrogen bonds, dipole-dipole forces London forces;
- know the three different but related bonding models(Lewis model, valence bond model and molecular orbital model) recognize the usefulness of the bonding theories in explaining and predicting molecular properties (bond angle, bond length, bond energy ,etc…);
- explain how the properties of a substance (solid or liquid) depends on the nature of the particles present the type of intermolecular forces;
- appreciate the importance of intermolecular forces in plant animal life:
- explain how the Valence Shell Electron Pair Repulsion(VSEPR) model can be used to predict molecular shape;
- know the types of crystalline solid (ionic, molecular, covalent network, or metallic) formed by a substance describe their properties;
- conduct experiments to observe analyze the physical properties of different substance to determine the type of bonding present;
- describe scientific enquiry skills along this unit: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, asking questions, applying concepts, relating cause and effect making generalization.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • Describe the reason why atoms form chemical bonds • State octet rule • Define chemical bonding	**3. Chemical Bonding Structure** **3.1. Introduction** (1 period) • Octet rule	Students should be able to define chemical bonding in the simplest way such as: i. that holds the parts of a chemical structure together or ii. a force which holds chemical species together. You might sometimes refer to bonding as a 'chemical bond', at other occasions refer to it as 'an interaction' or 'the force' holding molecules together. Yet different students may well have their own idiosyncratic ways of using these various terms that teachers are not aware of. All these terms which mean the same thing could be used interchangeably but students may not see it that way. Besides students should appreciate how bonding has resulted in the formation of an ever increasing number of substances from a limited number of elements substances with different properties from the elements they are formed substances accompanied with energy change(energy consuming or releasing)
• Describe the types of chemical bonding	• Types of chemical bonding	Students often believe that atoms will spontaneously form species with octet

Competencies	*Contents*	*Suggested Activities*
their mechanisms of the bonding processes		structures regardless of the chemical context. This could arise from the notion that species with octet electronic configurations or full outer shells are always more stable than species with other configurations. Such judgments are sometimes made regardless of the context any other factors that might be relevant. It is clear that the complete shell of electrons dominates students thinking even after looking at stability in terms of energy changes forces between charged particles. $^{-}$. The octet rule is an observation that correlates atomic stability with outer shell carrying eight electrons. Students should learn the types of chemical bonding how they relate to the electronic structures of the atoms that are bonded together. In addition to knowing what each of the bonds is, students should appreciate the similarities differences among them with especial emphasis to the covalent ionic ones. This of course should not down-play the importance of other types of bonding, especially in biological systems.
• Define ionic bonding.	**3.2 Ionic Bonding** (5 periods)	Students should know that an ionic compound is stable because of the very large lattice energy that is released when the oppositely charged ions come together to produce the ionic solid. If it were not for this lattice energy, ionic compounds would not exist. Thus in the gas phase the ions, Li^{+} (g) F^{-}(g), are less stable (i.e., of higher energy) than Li(g) F(g) therefore Li(g) F(g) would not react spontaneously to produce the ions. This is a good indication how easily students revert to simple familiar ideas. Students need to attempt chemical stability assignments individually discuss them in groups. Eg. Na^{+}/Na/Na^{7-}; Cl^{7+}/Cl/Cl^{-} C^{4+}/C/C^{4}.
• Use Lewis electron dot symbols to depict main group elements. • Describe ionic bonding using Lewis electron dot	• Lewis electron dot symbols	When discussing this section focus should be on: • Showing the structures of elements using the Lewis dot formulas as entry to showing what constitutes ionic bonds. • Formation set of criteria for the

Competencies	*Contents*	*Suggested Activities*
symbols. • List the favorable conditions for the formation of ionic bond. • Explain the formation of ionic bonding	• Formation of ionic bonding - The Born-Haber cycle	formation of ionic bonding • Appreciating how transfer of electrons between the atoms forming an ionic bond results in the formation of stable ions, isoelectronic to nearby noble gas elements • Illustrating the use of Born-Haber cycle
• Give examples of ionic compounds • Define lattice energy • Calculate lattice energy of ionic crystal from a given data using the Born-Haber cycle	- Factors affecting formation of ionic bonding	• Factors that influence the formation of ionic compounds need to be appreciated(one needs to be of low ionization energy electron affinity the second with high ionization energy electron affinity when the lattice energy of the resultant compound is high)
• Discuss the exceptions to octet rule	• Exceptions to octet rule	Students should be aware of the problem of over generalizing the Octet Rule; it works beautifully for period 2 elements but has limitations elsewhere
• Describe the properties ionic compounds • Carryout an activity to demonstrate the effect of electricity on ionic compounds (PbI_2 NaCl) • Carryout an activity to investigate the melting point solubility of some ionic compounds (NaCl $CuCl_2$)	• Properties of ionic compounds	Students should grasp the concept that properties of ionic compounds reflect the kinds of particles that are within them. Students should be aware that the strong attractive forces between ions cause ionic compounds to have their characteristic properties (it is important to highlight that ionic compounds conduct electricity when melted but not in the solid state) Students should be given samples of ionic compounds (PbI_2, NaCl, $CuCl_2$) asked to investigate their conductivity, solubility , melting points. Students should understand the high melting points are not exclusive to ionic substances. They need to be made aware that metals high melting points , as do macromolecular solids, the high boiling points in both cases result from the three dimensional attraction of each atom to its neighbors; this is the same factor occurs for ionic solids
• Define covalent bonding • Explain the formation of covalent bonding	**3.3 Covalent Bonding** (15 periods)	Students should be able to see that covalent bond is very different from ionic bond this in turn gives rise to very different properties from ionic compounds

Competencies	***Contents***	***Suggested Activities***
• Give examples of covalent molecules	• Formation of covalent bonding	sharing a pair of electrons between two atoms lowers their energy hence making the association more stable than the starting atoms covalent bonding is favored when the difference between the ionization energies electron affinities of the combining atoms is not large
• Draw Lewis structures or electron dot formulas of some covalent molecules	• Representation of covalent bond. - drawing Lewis structures	it is important from the outset to stress that substances containing covalent bonds only are called molecules with ample examples Students must see the connection between structure properties of molecules early on therefore must grasp the techniques of drawing Lewis structures of simple molecules. the need to determine skeletal structure be sure the students learn use the rules of counting distributing valence electrons among the bonds atoms in the skeletal structure
• Illustrate the formation of coordinate covalent bonding using examples	• Coordinate covalent bonding.	It is important to make clear to the students that the electrons to be shared by the two nuclei connected by a covalent bond need not be always contributed by both atoms; one could donate the second accept. This is how the Coordinate Covalent bond is formed. Difference in the source of the electrons may not result in difference in the covalent bonds formed. (eg the bonds in H_3O^+ are all identical even if it is formed from a reaction between H^+ H_2O)
• Define resonance structures	• Resonance structures	Students have to see the limitations of the Resonance concept see it only as a last resort to give a structure of a molecule when there is more than one reasonable way of distributing electron pairs in a molecule
• Draw resonance structures of some covalent molecules polyatomic ions.		
• Discuss the exceptions to the octet rule in covalent bonding	• Exceptions to the octet rule in covalent bonding	Students need to be made aware that there are exceptions to the octet rule by giving them few simple examples.
• Distinguish between polar non-polar	• Polar non-polar covalent	Students should be able to see that ionic substances are always polar (the extreme case of polarity is ionic bond) but that molecules could be either very polar, weakly polar or non-polar at all which dependence on the

Competencies	*Contents*	*Suggested Activities*
covalent molecules.	molecules	composition shape of the molecule. It is important that ample examples should be given here to determine the net dipole moments which ultimately decide polarity. Using the electro negativity table, students must predict bond stability (large electro negativity) dipole formation of molecule.
• Describe the properties of covalent molecules • Carryout an activity to investigate the effects of heat, electricity some solvents on covalent compounds (naphthalene, graphite, iodine ethanol)	• Properties of covalent compounds	Students should discuss general properties of covalent compounds. Students should grasp the concept that properties of covalent compounds reflect the kinds of particles that are within them. Students should appreciate how the properties of molecules depend on the net dipole moment in particular. Students should be given samples of covalent compounds (naphthalene, graphite, iodine ethanol) asked to investigate their conductivity solubility.
• Describe the valence shell electron pair repulsion theory (VSEPR) • Distinguish the bonding pairs non-bonding pairs of electrons. • Describe how electron pair arrangements of shapes of molecules can be predicted from the number of electron pairs. • Explain why double bonds lone pairs cause deviation from ideal bond angles. • Explain the term dipole moment with the help of a diagram	3.3.1Molecular Geometry • Valence shell electron pair repulsion (VSEPR) Theory • Electron pair arrangement molecular shape • Guidelines for applying VSEPR model. • Molecular shape molecular polarity - Bond polarity	Students should appreciate VSEPR theory its connection to Lewis diagrams of molecules in describing molecular shapes. Here it might be helpful to resemble Lewis structure with the blueprint of a building (which reveals part of the information) Molecular shape derived using VSEPR theory with the model of a building which is a miniature copy of the actual building that reveals vivid 3-D information in addition to what you get from the blueprint. Students can visualize the concept of lonepair-lonepair repulsion> lonepair-bondpair repulsion>bondpair-bondpair repulsion using a balloon model of water. Give assignments to the students individually to construct the five shapes of molecules discuss them in groups. Students should use guidelines for applying VSEPR theory to draw the shapes of molecules. Ball stick models are helpful to students as they attempt to visualize the geometry of molecules.

Competencies	*Contents*	*Suggested Activities*
• Describe the relationship between dipole moment molecular geometry • Describe how bond polarities molecular shapes combine to give molecular polarity. • Predict the geometrical shapes of some simple molecules • Construct models to represent shapes of some simple molecules.	- Bond angle - Dipole moment • Predicting the shapes of Molecules	Students should be given assignments to determine shapes polarity of molecules by combining the Lewis VSEPR theories. When considering bond polarity dipole moments should be treated qualitatively in conjunction to geometry of the bonds. Students should underst how some properties of similar compounds differ because of varying molecular shapes structures.
• Define intermolecular forces • Name the different types of intermolecular forces • Explain dipole-dipole interactions. • Give examples of dipole-dipole interaction • Define hydrogen bonding. • Explain the effects of hydrogen bond on the properties of substances • Give reason why H-bonding is stronger than ordinary dipole-dipole interaction.	3.3.2. Inter molecular forces in covalent compounds • Dipole- dipole force -hydrogen bonding	Students should be aware that the ability to predict avoids the necessity of memorizing hundreds of individual molecular shapes. Students should be asked to construct models to represent shapes of simple molecules. It is very important to help students see that the particles in molecular substances are molecules whereas in ionic substances the particles are positive negative ions. The fact that molecular substances are either gases, liquids or soft solids whereas ionic substances are usually hard solids with high melting point does not imply that covalent bonds are in general weaker than ionic bonds; the forces that should be considered in molecular substances is the various types of intermolecular associations or bonds with varying strength. Students often do not count intermolecular forces as bonds these needs to be corrected. Students should know that an intermolecular force is a force of attraction between molecules. Students be aware that Dipole-Dipole force hydrogen bonding occur in some molecules, but Dispersion or London forces are common to all molecules. Students should use polarity to explain

Competencies	*Contents*	*Suggested Activities*
• Explain dispersion (London) forces • Give examples of dispersion forces • Predict the strength of inter molecular forces for a given pair of molecules.	• Dispersion or London force	hydrogen bonding should understand the properties of substances having hydrogen bonding. It is important to help that students recognize that the hydrogen bonding that causes ice to occupy more volume therefore less dense is important both to ice skaters to aquatic plants animal. If ice were denser, the ponds streams would freeze from the bottom up. The plants animals would die. Students should know that the weak attractions due to instantaneous induced dipoles were described first in 1930s by the German physicist Fritz London, so they are called London forces. Students should be aware that substances with molecules that are attracted to one another only by London forces have low boiling low freezing points, compared with other substances of about the same molecular weight. If their molecules are small they usually exist as gases at room temperature
• Explains how metallic bond is formed. • Explain the properties of metals related to the concept of bonding. • Carryout an activity to investigate the conductivity, malleability ductility of some metals nonmetals(Al, Cu, Fe, Sn, Zn, S, C- charcoal, C- graphite Si)	**3.4. Metallic Bonding** (2 periods) • Formation of metallic bond • Electron Sea Model • Properties of metals related to the concept of bonding.	If all chemical bonding is understood only in terms of striving to obtain a full outer shell by sharing (covalent bonding) or transferring (ionic bonding) electrons, then we run into problem when we consider the force that holds metallic atoms together. It might help here to tell the students that it is not a model as easy as the covalent or ionic ones. Metallic bonding is not the sharing or loss or gain of electrons. It is just a loose association of metal ions electrons they have lost. Students should know that the electron Sea model depicts metallic bonding as a situation where all the metal atoms in a sample pool their valence electrons into an evenly distributed 'sea' of electrons that 'flows' between around the metal-ion cores attracts them together. Unlike the localized electrons in covalent bonding, electrons in metallic bonding are delocalized, moving freely throughout the piece of metal. Students should be aware that the physical properties of metals vary over a wide range. Almost all are solids with moderate to very

Competencies	*Contents*	*Suggested Activities*
		high melting extremely high boiling points. Periodic trends are consistent with the model. Ca has higher melting point than K- a reflection of 2 vs 1 valence electrons. They are malleable ductile-the metal ions slide past each other when force is applied. They are good conductors of electricity heat-because of the mobility of the delocalized electrons.
• Name the two chemical bond theories. • Explain the valence bond theory. • Distinguish between the Lewis model the valence bond model. • Discuss the overlapping of orbital in covalent bond formation • Explain hybridization • Show the process of hybridization involved in some covalent molecules • Draw hybridization diagram for the formation of sp, sp^2, sp^3, sp^3d sp^3d^2 hybrids. • Suggest the kind of hybrid orbitals on the basis of the electron structure of the central atom. • Predict the geometrical shapes of some simple molecules on the bases of hybridization the nature of electron pairs. • Discuss the hybridization involved in compounds containing multiple bonds.	**3.5 Chemical Bonding Theories** (8 periods) 3.5.1.Valence Bond Theory (VBT) - Overlap of atomic orbitals (sigma Pi bonds) - Hybridization of orbitals	Students could be given samples of metals non-metals (Aluminum, copper, iron, tin, zinc, charcoal, graphite silicon) asked to investigate their conductivity , malleability ductility. Extra care has to be taken to present that both MOT VBT are complementary therefore should be used when one has more advantage than the other. Students could get frustrated as they would like one that works best all the time. This section describes the modern theories of chemical bonding how they can account for (or predict, in some cases) molecular structure. The thing to keep in mind throughout discussions of the different approaches is that each theory is attempting to describe the same thing each, in its own way, succeeds to a degree. The theories, then, present alternative views of the same phenomenon. Students have to be aware that they complement one another; what VBT may not be conversant with could be explained by MOT vice versa. Students should be aware that VBT is a theory that explains covalent bond formation by hybridization overlapping of atomic orbitals. The concept of orbital hybridization is so involved it has to be kept to the minimum possible. Students could draw hybridization diagram predict the geometrical shapes of some simple molecules.

Competencies	*Contents*	*Suggested Activities*
• Explain bond strength bond length.		
• Explain the molecular orbital theory. • Describe molecular orbital using atomic orbitals. • Describe bonding anti-bonding orbitals.	3.5.2.Molecular Orbital Theory (MOT) • Combination of atomic orbitals • Bonding antibonding molecular orbital	This section presents an elementary treatment of the molecular orbital theory. Students are certain to encounter this theory in more advanced courses. Students should be aware that MOT is a theory of the electron structure of molecules in terms of molecular orbitals , which may spread atoms or the entire molecule. Students should learn the differences between bonding antibonding molecular orbitals how the electronic structure of a molecule is obtained by filling molecular orbitals. They should also see that MOT avoids the idea of resonance which is part parcel of VBT. The discussion of how molecular orbitals are formed from electron waves of atomic orbitals has to be qualitative introductory. Show the students the similarity between atomic orbitals in atoms molecular orbitals in molecules regarding filling with electrons. Construct molecular orbitals for diatomic molecules such as H_2, He_2,N_2 O_2.
• Draw molecular orbital energy level diagrams for mononuclear diatomic molecules. • Write the electron configuration of simple molecules using molecular orbital model.	• Electron configuration of diatomic molecule	Students should learn how to calculate the net bond orders from the number of bonding antibonding electrons. It is wise to see the similarity with that of the Lewis formula.
• Define bond order determine the bond order of some simple molecules ions • Determine the stability of a molecule or an ion using its bond order. • Predict magnetic properties of molecule.	• Bond order	Students should note that the bond order can only be used qualitatively for purpose of comparison. For example, two electrons in a sigma molecular orbital or two electrons in a bonding pi orbital would have a bond order of one. Yet these two bonds must differ in bond strength bond length because of the difference extent of overlap of atomic orbitals. Give some examples to show the strengths of MOT (eg magnetism).

Competencies	*Contents*	*Suggested Activities*
	- Magnetic properties	
- Define crystal - Name the four types of crystalline solids give examples. - Mention the type of attractive forces that exist within each type of crystalline solids - Describe the properties of each type of crystalline solids - Build a model of sodium chloride crystal structure	**3.6.Types of crystals (1 period)** - Ionic crystal - Molecular crystal - Covalent network crystal - Metallic crystal	Students should know that types of crystals are distinguished from each other by the type of intermolecular forces between the individual molecular particles or ions which in turn determine physical properties like melting boiling points solubility in different solvents. Students should be asked to construct a model of NaCl structure from ball sticks or other locally available materials.

Assessment

The teacher should assess each student's work continuously over the whole unit 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: Describe the reason why atoms form chemical bonds; define chemical bonding, ionic bonding, lattice energy, covalent bonding, resonance structures, intermolecular force, hydrogen bonding and metallic bonding; describe the types of chemical bonding and their mechanisms of the bonding processes; explain the formation of ionic bonding and covalent bonding, and give examples of ionic and covalent compounds; calculate lattice energy of ionic crystal from a given data using Born-Haber-cycle; discuss the exceptions to Octet rule in ionic and covalent bonding; describe ionic bonding using Lewis electron dot symbols and the Lewis properties of ionic compounds; draw Lewis structures and resonance structures for some covalent molecules and polyatomic ions; illustrate the formation of coordinate covalent bonding using examples; describe the properties of covalent molecules; distinguished between polar and non-polar molecules; describe the VSEPR theory and how electron pair arrangements of molecules can be predicted from the number of electron pairs; define the bonding and non-bonding pairs of electrons; explain dipole moment describe the relationship between dipole moment and molecular geometry; predict the geometrical shapes of some simple molecules; explain that the VSEPR theory is used to explain the molecular parameters like bond angle and polarity of the molecule; construct models to represent shapes of some simple molecules; name the different types of intermolecular forces; explain dipole-dipole interactions and dispersion forces and give examples of them; explain the effect of hydrogen bonding; explain the valence bond and molecular orbital theories; discuss the overlapping of orbitals in covalent bond formation; explain hybridization, show its process, draw hybridization diagrams for the formation of hybrids and discuss the hybridization

involving multiple bonds; describe molecular orbitals; write the bonding antibonding orbitals; write the electron configuration of simple molecules using MOT; define bond order and determine the bond order of same simple molecules and ions; determine the stability of a molecule or an ion using its bond order; explain the properties of metals;
define crystal, name the four types of crystals and give examples mention the type of attractive forces, and properties of each types of crystals; build a model of Nacl crystal structure;

Students above minimum requirement level

Students working above the minimum requirement level should be praised their achievements recognized. They should be encouraged to continue working hard 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 additional lesson time during breaks or at the end of the day.

Unit 4: Chemical Kinetics *(19 periods)*

Unit outcomes: Students will able to:

- explain what is meant by reaction rate perform the activities to determine it;
- demonstrate an understanding of the dependence of reaction rate on nature of reactants , surface area of solid reactants, concentration of reactants or pressure of a system, temperature of a system catalysts;
- determine the reaction rates using experimental data calculations;
- describe how reaction rate theories (collision theory transition state theory) can be used to explain changes in reaction rates;
- determine the rate laws order of reactions from data on initial concentrations reaction rates;
- understand that most reactions occur as a series of elementary steps in reaction mechanisms;
- appreciate the importance of chemical kinetics in industry in living organisms;
- describe scientific enquiry skills along this unit: observing, predicting, comparing and contrasting, communicating, asking questions, drawing conclusion, applying concepts, relating cause and effect, making generalization problem solving.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • define chemical kinetics; • define reaction rate; • express the reaction rate in terms of the disappearance of the reactants the appearance of the products;	**4. Chemical kinetics** **4.1 Rate of Reaction** (5 periods)	Students should be able to see how applied this topic is in their daily life as well as in many other scientific disciplines. Students should appreciate that chemical kinetics is the study of the rates mechanisms of chemical reactions. Iron rusts in moist air faster than in dry air; foods spoil faster when left unrefregerated; one's skin becomes tanned more readily in the summer than in the winter. These are three common examples of complex chemical changes with rates that vary according to the reaction condition. To enhance rates of chemical reactions whose products are valuable, or to slow down rates of chemical reactions whose products are harmful, one has to understand how chemical rates are affected. Knowledge of reaction rates of chemical reactions can also be used in conjunction with factors that influence them to get insight into the sequence of chemical steps that occurs to produce the overall net reaction. Students should be aware that the reaction rate is expressed as a change in the concentration of reactants or products per unit time

Competencies	*Contents*	*Suggested Activities*
• perform an activity to measure rate of reaction; • calculate the reaction rate from a given data	• Determination of reaction rate	Students could perform an activity to measure the rate of reaction between marble chips($CaCO_3$) dilute HCl acid. Students should be given the chance to generate rate data from experimental as well as hypothetical change of concentration of reactants/products in time. Students should work in groups individually to plot graphs calculate rates by determining slopes at different periods of time.
• list the factors that affect reaction rate • explain how nature of reactants, surface area of solid reactants, temperature concentration or pressure affect the rate of reaction by giving examples • explain activation energy • define catalyst, positive catalyst negative catalyst • distinguish between homogeneous heterogeneous catalysis • distinguish between biological non-biological catalysts • draw energy diagram that represents activation energy show the effect of catalysts • distinguish between homogeneous heterogeneous reactions • perform an activity to show the effect of nature of reactants on reaction rate • perform an activity to show effect of surface area on reaction rate	• Factors that affect the reaction rate	The factors that affect rates of chemical reactions should be given due coverage with examples. Students need to see how rates of chemical reactions change with temperature, concentration, nature of the reactant, surface area availability of a catalyst. • The temperature of the system is important because all chemical reaction rates increase markedly with increasing temperature • The concentrations of the reactants continually decrease as the reaction progresses; in most cases this also means that the instantaneous rate of reaction decreases with time • The nature of reactants can account for enormous differences in reaction rates • The increased surface area of solid reactants increaser the rate of reaction • A catalyst can increase the reaction rate by forming an intermediate compound with a relatively low activation energy for a reaction (Homogeneous Catalysis), or by absorbing reacting species on its surface(Heterogeneous Catalysis) Students can be assigned one concentration each then pool their data together to see the effect. This could be applied to the other factors as well to

Competencies	*Contents*	*Suggested Activities*
• perform an activity to show the effect of concentration on reaction rate • perform an activity to show the effect of temperature on reaction rate • Perform an activity to show the effect of catalysts on reaction rate		minimize the amount of time it could possibly take.
		Students can perform activities to show the effect of Nature of reactants on reaction rate i) $Cu + HCl \longrightarrow$ ii) $Mg + HCl \longrightarrow$ Surface area of reactants on reaction rate $CaCO_3$ (cube chalk) + HCl (dilute) $\longrightarrow$ $CaCO_3$ (powderedchalk) + HCl(dilute) $\longrightarrow$ Concentration of reactants on reaction rate $CaCO_3$ + 1M HCl $\longrightarrow$ $CaCO_3$ + 2M HCl $\longrightarrow$ Temperature on reaction rate i) $H_2SO_3 + NaS_2O_3 \longrightarrow$ (At low temperature) ii) $H_2SO_3 + NaS_2O_3 \longrightarrow$ (At high temperature) Catalyst on reaction rate (Decomposition of hydrogen peroxide, H_2O_2) i) H_2O_2(without catalyst) $\longrightarrow$ $H_2O + O_2$(amount ?) ii) H_2O_2(with catalyst, MnO_2) $\longrightarrow$ $H_2O + O_2$(amount ?)
	4.2 Theories of Reaction Rates (4 periods)	Students should know that there are two theories running parallel that attempt to explain how rates of reactions are affected by the • nature of the reactants products • concentration of the reactants • temperature • influence of catalysts • surfaces area of solid reactants

Competencies	*Contents*	*Suggested Activities*
• state collision theory • define activation energy' • describe how collision theory can be used to explain changes in reaction rate;	• Collision theory - Concentration collision theory	The two theories are distinct but in complete agreement, each emphasizing on different aspects of the reaction process. Students should know that collision theory views the reaction rate as the result of particles colliding with a certain frequency minimum energy. The basis for the collision theory is the notion that molecules must collide in order to react. We can predict the rate law if we know what collisions take place during a reaction.. Factors that increase collision frequency, hence, increase the rate of the reaction. As a corollary to this, students should know that if there is no collision at all then there is no reaction at all its rate becomes zero. Out of the factors that affect rates mentioned above, concentration temperature have direct relevance in this theory. Students could see how collision frequency increases if concentration is increased.
	- Temperature collision theory	Students should be aware that hi temperature implies higher kinetic en from physics. Remind students that K.E = ½ mv^2, w implies that the molecules will co frequently effectively which in turn implies higher The minimum kinetic energy req between two colliding molecules in ord produce an effective collision is calle activation energy, E_a.
		Students should know that transition theory views chemical kinetic of energy geometry of the activated complex that it has formed come apart to yield reac again go to produce the products .
• state transition state theory • define activated complex (transition state) • describe how transition	• Transition state theory	Students should draw label energy diag for both exothermic endothermic reac

Competencies	*Contents*	*Suggested Activities*
state theory can be used to explain changes in reaction rate;		with the terms transition state, acti complex, activation energy barrier fo forward reverse reactions heat of reacti
• sketch label the energy profiles of reactions which are exothermic endothermic	• Energy profiles of exothermic endothermic reaction.	
• define rate law, order of reaction rate constant • determine reaction order • calculate rate constants from a given experimental data • describe the role of the rate constant in the theoretical determination of a reaction rate	**4.3 Rate Equation or Rate Law** (7 periods) • Order of reaction • Rate constant	Students should understand the effect of concentration of reactants on rates of reaction by using rate law. The rate law for a reaction such as: $A + B \rightarrow$ Products is Rate = $k[A]^x [B]^y$ where k is the rate constant. The exponents, x y, are called the order of the reaction with respect to A B, respectively. The overall order is (x + y). Emphasize the fact that the actual values of x y can only be obtained by experiment. They have no connection to the coefficients of the balanced chemical equation. If they happen to be the same then it is mere chance. These experiments involve observing what effect altering the concentrations of the reactants has on the rate of reaction. For a reaction, $A \rightarrow B$, i. if doubling the concentration of A doubles the rate, the exponent is 1. Rate = $k[A]^1$ ii. if doubling the concentration of A quadruples the rate, the exponent is 2. Rate = $k[A]^2$. Students should be given mixed exercises so that they are able to see orders with values zero fractions. The rate constant is the proportionality constant, k, in the rate law given above. Its value is the rate of the reaction when all the reactant concentrations are 1M. The rate constant is the one variable that gives the information about the nature of the reactants. According to Arrhenius $k = Ae^{-Ea/RT}$, A is related about the shape overall nature of the reactant, Ea is the activation energy, T is the temperature

Competencies	*Contents*	*Suggested Activities*
		in Kelvin, R is the universal gas constant e is base of natural logarithm.. Students should be guided to see how big k becomes when A T are big Ea low vice versa. Discussion should be limited to qualitative explanations.
• explain the zero, first second order reactions using concentrations versus time curve • calculate that concentration time of the reaction mixture for different order reactions plot their graphs • define half-life • calculate the half-lives of zero order, first order second order reactions from experimental data	• Concentration - Time Equation • Half - life of a reaction	A rate law tells us the rate of reaction depends on reactant concentrations at a particular moment. But students should be aware that there is another mathematical relationship which shows how a reactant concentration changes over a period of time; that is Concentration-Time equation. Students should know that the concentration of the reactant decreases as a reaction proceeds because it is being consumed. The time that takes for the reactant concentration to decrease to one-half of its initial value is called the half-life of a reaction. Students could determine the half-life o first the second order reactions.
• explain reaction mechanism • explain molecularity of a reaction • explain what is meant by rate determining step • give specific examples to illustrate rate determining step	**4.4 Reaction Mechanism** (3 periods) • Molecularity of an elementary reaction - Unimolecular - Bimolecular - Termolecular	Students should be aware that one of the objectives of studying reaction rates of chemical reactions is to get an insight into the sequence of chemical steps (reaction mechanism) that occurs to produce the overall net reaction. The individual reactions that make up a mechanism are called elementary steps. An elementary step is one of the individual steps in a reaction mechanism, which can be either • **unimolecular**: involving the decomposition or rearrangement of a single particle • **bimolecular**: involving collision of two particles • **termolecular**: involving collision of three particles all at one time Higher molecularities are not believed to take place because of low probability. When we have to choose

Competencies	*Contents*	*Suggested Activities*
• explain the relationship between the reaction path way the rate law • use rate equation to suggest possible reaction mechanism for a reaction	• Rate determining step	between alternative mechanisms for a reaction the one with lower molecularity are preferred, all other things being equal. Obtaining a satisfactory mechanism for a reaction is a very difficult task even for experienced chemists; therefore this has to be limited only to qualitative treatment. Students should be aware that when there are more than one step in a mechanism of a reaction, the slowest step is called the rate-determining step because the final products cannot be formed any faster than the products of the slowest step. Students should know that the predicted rate law only includes the reactants in the overall equation. Neither intermediates nor catalysts appear in the rate law.

Assessment

The teacher should assess each student's work continuously over the whole unit 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 kinetics and reactions rate; express the reaction rate in terms of the disappearance of reactants and the appearance of the products calculate the reaction rate from a given data; perform an activity to measure rate of reaction; list the factors that affect reaction rate and explain how they affect using examples; define catalyst, positive catalyst negative catalyst; distinguish between homogeneous and heterogeneous catalysts, biological non-biological catalysts, and homogeneous and heterogeneous reactions; perform activities to show the effect of nature of reactants, surface area of solid reactants, concentration, temperature catalysts on rate of reaction; state collision theory and describe how it can be used to explain changes in reaction rate; state transition state theory and describe how it can be used to explain changes in reaction rate; define activation energy, activated complex, rate law, order of reaction and rate constant; sketch label the energy profiles of reactions which are exothermic endothermic; determine reaction order and calculate rate constants from a given experimental data; explain the zero, first second order reactions using concentrations verses time curve; define half-life calculate the half-lives of zero, first second orders reactions from experimental data; explain reaction mechanism molecularity of a reaction; explain rate determining step give examples to illustrate it; explain the path way and the rate law use rate equation to suggest possible reaction mechanism for a reaction;

Students above minimum requirement level

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Students below minimum requirement level

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Unit 5: Chemical Equilibrium Phase Equilibrium *(18 periods)*

Unit outcomes: Students will be able to:

- understand how equilibrium is established;
- explain characteristics of dynamic equilibrium;
- state the law of mass action write an expression for equilibrium constants, K_c K_p, from a given chemical reaction;
- apply the law of mass action to calculate K_c , K_p, concentration pressure of substances in equilibrium;
- understand how the reaction quotient is used to indicate position of the equilibrium;
- state Le-Chatliers' principle use it to predict explain the effects of changes in temperature, pressure, concentration presence of a catalyst on a reaction;
- perform an activity to demonstrate the effects of changes in concentration on the position of equilibrium to determine Kc or K_p values;
- explain how equilibrium principles may be applied to optimize the production of industrial chemicals (e.g. production of ammonia sulphuric acid);
- describe scientific enquiry skills along this unit: observing, predicting, comparing and contrasting, communicating, asking questions making generalization.

Competencies	*Contents*	*Suggested Activities*
Students will be able to: • explain reversible irreversible reactions • define dynamic chemical equilibrium • state the necessary conditions for attainment of equilibrium • describe the microscopic event that occur when a chemical system is in equilibrium • explain the characteristics of chemical equilibrium	**5. Chemical Equilibrium phase equilibrium** **5.1 Chemical equilibrium** (14 periods) • Reversible irreversible reactions • Attainment characteristics of chemical equilibrium	Since students may form a misconception of chemical equilibrium easily while thinking it is simple, enough explanation of the definition with the help of examples needs to be given from the beginning. Students should appreciate that chemical equilibrium applies only to reversible chemical reactions is a dynamic one (not static), that what remain unchanged in a chemical reaction at equilibrium are the rates of both the forward reverse reactions that they are equal; hence no shift in position of equilibrium, concentrations of reactants concentration of products remain the same (but concentration of reactants are not necessarily equal to concentration of products; if they are it is mere coincidence)

Competencies	***Contents***	***Suggested Activities***
		Students should be aware that dynamic chemical equilibrium has three properties. They are:- ▪ Dynamic chemical equilibrium can be reached from either direction ▪ Energy change at dynamic chemical equilibrium is zero ▪ Reactants products have different quantities than were present initially
• state the law of mass action	▪ Equilibrium expression equilibrium constant - The law of mass action	Students should know that there is a simple relationship between the concentrations of reactants products in an equilibrium system. The relationship between mass action expression, reaction quotient equilibrium constant has to be appreciated by students. They must be able to see that the equilibrium constant is a special reaction quotient which itself is a numerical value of mass action expression that is determined when the system is at equilibrium. The use of reaction quotient in relation to the equilibrium constant, as an indicator which way a reaction will go to reach equilibrium, has to be stressed. Applications of the equilibrium constant in determining the composition of the equilibrium mixture needs highlighting.
• define equilibrium constant • write the equilibrium constant expression for chemical reactions that involve concentration	- Equilibrium constant expression involving concentration, K_C	At equilibrium, at a given temperature, the mass action expression for a given reaction is always equal to the same number, the equilibrium constant. Remember, there are no restrictions on the individual equilibrium concentrations. The only requirement is that when they are substituted into the mass action expression, the resulting fraction must equal the

Competencies	***Contents***	***Suggested Activities***
• calculate values for equilibrium constant involving concentration • state the relationship of K_{eq} to the relative amounts of products reactants in a given reaction	- Equilibrium constant expression involving partial pressure, K_P	equilibrium constant. For the general reaction aA + bB ↔ mM + nN the mass action expression is: $K_c = \frac{[M]^m[N]^n}{[A]^a[B]^b}$ Where Kc is equilibrium constant, when all concentrations are equilibrium concentrations.
• write the equilibrium constant expression for chemical reactions that involve partial pressure; • calculate values for equilibrium constant involving partial pressure;		Students should note that for reactions involving gases, partial pressures can be used in place of molar concentrations. Remember the relationship $K_p = K_c(RT)^{\Delta n}$ which is used to interchange K_p K_c.. Δn is change in the number of moles of gases.
• show the relationship between K_C K_P	- Equilibrium constant for heterogeneous reaction	Students should underst that in heterogeneous reactions the reactants products are not all in the same phase. The concentrations of pure liquid solid are not included in the mass action expression of the heterogeneous reactions. This is because they are constant are included in the equilibrium constant.
• distinguish between homogeneous heterogeneous reactions; • define reaction quotient • use the equilibrium quotient to predict the direction of the reaction the position of equilibrium • calculate equilibrium concentrations given initial concentrations • determine whether the reactants or products are	- Reaction quotient (Q_C) - Predicting the direction of reaction - Applications of equilibrium constant • Predicting the position of equilibrium mixture	Students should know that reaction quotient, Q_C, is the numerical value of the mass action. It is used to predict whether a reaction has reached or not if it has not which direction it will proceed. If $K_c = Q_C$, system is at equilibrium, $K_c < Q_C$, reaction proceeds toward reactants $K_c > Q_C$, reaction proceeds toward products. There are basically two kinds of calculations that students have to

Competencies	*Contents*	*Suggested Activities*
favored in a chemical reaction given the equilibrium constant;	• Calculating the equilibrium concentrations of reactants products	learn. One is to calculate K from either equilibrium concentrations or information from which you can deduce equilibrium concentrations. The other is to calculate information about equilibrium concentrations, given the value of K. Students should be given exercises until they become conversant. Students should underst the concepts behind the mathematical manipulation of equilibrium problems.
• list factors that affect chemical equilibrium • state Le-chatliers principle • use le-chatliers principle to explain the effect of changes in temperature, pressure, concentration presence of catalyst on a reaction; • state the effect of changes in concentration, pressure/volume temperature on K_{eq} • perform an activity to demonstrate the effect of changes in concentration on the	▪ Changing equilibrium conditions: Le- chatleries principle	One important application of this topic is described as the LeChatelliers principle. As much as it is useful it is easily confused by students. It needs to be defined clearly in simple language. One common point confused here is that after new equilibrium is established right after subjecting a system in equilibrium to stress , the equilibrium constant remains unchanged but the individual concentrations of reactants products could change. The point that, equilibrium constant for a given reaction changes only with temperature but not with concentration or pressure of reactants/products has to be appreciated by students. Students should be able to see the application of LeChatelier's principle in industrial production of important substances such as ammonia sulfuric acid. ▪ When a reactant or product is added to a system at equilibrium, the position of equilibrium shifts toward the

Competencies	*Contents*	*Suggested Activities*
position of equilibrium • perform an activity to demonstrate the effect of changes in temperature on the position of equilibrium • Perform an activity to determine K_C for esterification of an organic acid		opposite side of the equation. ▪ Decreasing the concentration of a reactant or product causes the position of equilibrium to shift in the direction of the substance removed. ▪ Increasing the pressure by decreasing the volume shifts the position of equilibrium in the direction of the fewest number of moles of gas. ▪ An increase in temperature causes the position of equilibrium to shift in the direction of the endothermic reaction. ▪ A decrease in temperature causes the position of equilibrium to shift in the direction of the exothermic reaction. ▪ A very important point to remember is that the only thing that changes K for a reaction is a change in temperature!! ▪ Adding an inert (unreactive) gas to a system, without changing the volume, has no effect on the position of equilibrium. ▪ A catalyst has no effect on the position of equilibrium. It only increases the speed with which the system reaches equilibrium. Students could perform activities to demonstrate the effects of changes -In concentration on the position of equilibrium. Eg. Take 0.01M, 0.02M, 0.03M 0.04M $Fe(NO_3)_3$ (aq) to react with 0.01M KSCN(aq) -In temperature on the position of equilibrium. Eg. The reaction between iodine, I_2 starch

Competencies	*Contents*	*Suggested Activities*
		Students could perform activities to determine equilibrium constant. Eg. Kc for esterification of an organic acid
• define optimum conditions • explain how Le-Chatliers principle is applied	▪ Equilibrium Industry - Equilibrium in the Haber process - Equilibrium in the contact process	Students should know that chemical equilibrium along with LeChatelier's principle has applications in industrial production of important substances. In order to optimize production of ammonia in the Haber process sulfuric acid in the contact process, chemists or chemical engineers exploit fully the principles in this unit.
- in the Haber process (production of NH_3) - in the Contact process (production of H_2SO_4) • explain the interplay between kinetic potential energy underlies the properties of the three states of matter their phase changes • explain the process involved, both within a phase through a phase change when heat is added or removed from a pure substance • explain the meaning of vapor pressure how phase changes are dynamic equilibrium processes; • explain the relation between vapor pressure boiling point; • describe how a phase diagram shows a phase of	**5.2 Phase equilibrium (4 periods)** 5.2.1 Common terms - phase components, degree of freedom 5.2.2 Phase rule phase diagram - Phase diagram of water	This topic is a difficult one. Hence its coverage should be limited to introducing the terms: phase, phase rule, degree of freedom, etc. You could ask students to draw label phase diagrams for water CO_2 to show the triple point critical point.

Competencies	*Contents*	*Suggested Activities*
a substance at differing conditions of pressure temperature; • use a phase diagram, determine melting point, boiling point, critical temperature, critical pressure triple point of a substance		

Assessment

The teacher should assess each student's work continuously over the whole unit 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: Explain reversible and irreversible reactions; define dynamic equilibrium and state the necessary conditions for attainment of equilibrium; explain the characteristics of chemical equilibrium; state the law of mass action and define equilibrium constant and reaction quotient; write the equilibrium constant expression for chemical reaction that involve concentration and partial pressure, and calculate the values for equilibrium constant; determine whether the reactants or products are favored in a chemical reaction given the equilibrium constant; distinguish between homogeneous and heterogeneous reactions; show the relationship between K_C and K_P; use the equilibrium quotient to predict the direction of the reaction and the position of equilibrium; calculate equilibrium concentrations of reactants and products;
list factors that affect chemical equilibrium; state Le-chatlier's principle, use it to explain the effect of changes in temperature, pressure, concentration and presence of catalyst on a reaction and perform activities to demonstrate them; explain how Le-chatlier's principle is applied in the Haber and Ostwald's processes; explain the properties of the three states of matter and their phase changes; explain vapor pressure and the relation between vapor pressure and phase changes; describe how phase changes are dynamic equilibrium processes; describe how a phase diagram shows a phase of a substance at differing conditions of pressure and temperature;

Students above minimum requirement level

Students working above the minimum requirement level should be praised their achievements recognized. They should be encouraged to continue working hard 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 additional lesson time during breaks or at the end of the day.

Unit 6: Carboxylic Acids, Esters, Fats Oils *(20 periods)*

Unit outcomes: Students will be able to:

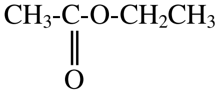
- write the structural formula IUPAC names of given carboxylic acids esters;
- describe some physical chemical properties of carboxylic acids esters;
- predict correctly name the products of organic reactions, including substitution, addition, elimination, esterification, hydrolysis oxidation reactions;
- carry out activities to prepare a carboxylic acid an ester;
- list some important fatty acids;
- test for the carboxylic acid ester functional groups;
- understand the structures, properties uses of fats oils;
- describe explain reactions in soap making (saponification);
- describe explain the cleaning action of soap detergents;
- understand the harmful effects of detergents on the environment when improperly disposed;
- describe scientific enquiry skills along this unit: observing, classifying, comparing and contrasting, asking questions, drawing conclusion, applying concepts problem solving.

Competencies	***Contents***	***Suggested Activities***
Students will be able to: • List common organic acids mention their sources; • Write the general formula of saturated monocarboxylic acids; • Write the molecular formulas names of the first six members of saturated monocarboxylic acids; • Give the structural formula for the first four members of saturated mono carboxylic acids. • Give examples for mono carboxylic, dicarbcxylic tricarboxylic acids; • Name some branched carboxylic acids;	**6 Carboxylic acids, Esters, Fats Oils** **6.1 Carboxylic acids** (8 periods) • Nomenclature structure	This unit deals with a narrow range of oxygen containing hydrocarbons that are abundant in nature, especially in living organisms. Students should know that carboxylic acids contain the carboxyl group attached to hydrogen (HCOOH), an alkyl group (RCOOH), or an aryl group (ArCOOH). The aliphatic carboxylic acids have been known for a long time, as a result have common names that refer to their sources rather than to their chemical structures. You could motivate the students if you relate common names of some carboxylic acids to the source they were isolated from initially. Example is butyric acid from rancid butter. You should not spend more time than is necessary on trivial or common names. As the number of both natural synthetic organic compounds, carboxylic acid included, has reached thouss, is still on the rise by the day; one cannot rely on common names. Students should appreciate the need for a nomenclature system that is systematic, easy to use, remember. This could then be followed by the IUPAC nomenclature system lesson. Even though IUPAC should be at the center of nomenclature, it is useful to

Competencies	*Contents*	*Suggested Activities*
		introduce the α, β, γ, δ…position of attachment. Direct the students attention to the fact that the carboxyl carbon is always considered as C-1, hence C-2 corresponds to α of the common names, C-3 to β, so on. (Caution: Do not mix Greek letters with IUPAC names, or Arabic numerals with common names.) Hence the α-carbon is the one bearing the carboxyl group. This is commonly used in biochemical literature. γ β α C-C-C-COOH (used in common names) Students should be able to discern the structural differences between the aromatic carboxylic acids aliphatic acids with phenyl ring substituents. The aromatic acids are usually named as derivatives of the parent acid, benzoic acid, C_6H_5COOH. Since the IUPAC names for the acids follow the usual pattern, you could give assignments to the students to see if they still remember what they learned earlier about organic nomenclature. Ask the students if they can point out the parent structures correctly give accurate names.
• describe the physical properties of saturated mono carboxylic acids • explain why "tella" or "tej" turns sour	• Physical properties	Ask students to list down some physical properties using the chemical knowledge they have mastered so far. You might hint that properties of molecules follow structure, to the students at this point see how this might help them. **Solubility**: Carboxylic acids have two distinct parts in terms of associating with water. Ask the students to identify them. They could discuss this in groups. The hydrocarbon part is hydrophobic whereas the carboxyl group is hydrophilic. The aliphatic acids, therefore, show very much the same solubility behavior as the alcohols: the first four or five are miscible with water the higher acids are virtually insoluble. Water

Competencies	*Contents*	*Suggested Activities*
		solubility undoubtedly arises from hydrogen bonding between the carboxylic acid water. The simplest aromatic acid, benzoic acid, contains too many carbon atoms to show appreciable solubility in water. Carboxylic acids are soluble in less polar solvents like ether, alcohol, benzene, etc. **Boiling Point**: As a class the carboxylic acids are even higher boiling than alcohols. For example propionic acid (b.p. 141 ◦C) boils more than 20 ◦C higher than the alcohol of comparable molecular weight, n-butyl alcohol(b.p. 118 ◦C). These very high boiling points are due to the fact that a pair of carboxylic acid molecules are held together not by one but by two hydrogen bonds. Ask your students to see if they can draw structure of two carboxylic acids with two intermolecular H-bonds between them. **Odor**: Ask the students to list down odors of some carboxylic acids that they might have encountered in laboratories or elsewhere. (Caution: Students are not allowed to inhale vapors of chemicals; if it is important that they smell chemicals it has to be supervised by the instructor they should only blow the vapor with their h in the direction of their noses.) The odors of the lower aliphatic acids progress from the sharp, irritating odors of methanoic acid ethanoic acids to the distinctly unpleasant odors of butanoic, pentanoic hexanoic acids; the higher acids have little odor because of their low volatility.
• explain the chemical properties of saturated mono carboxylic acids	• Chemical properties	Students should be aware that the characteristic chemical behavior of carboxylic acids is, of course, determined by their functional group, carboxyl, -COOH. This group is made up of a carbonyl group (C=O) a hydroxyl group (-OH). It is the –OH that actually undergoes nearly every

Competencies	*Contents*	*Suggested Activities*
		reaction-loss of H^+, or replacement by another group-but it does so in a way that is possible only because of the effect of the C=O. The most important chemical property of carboxylic acids is their acidity. They form salts when treated with strong bases like OH^- even with weaker bases such as HCO_3^- NH_3. The rest of the molecule undergoes reactions characteristic of its structure; it may be aliphatic or aromatic, saturated or unsaturated, may contain a variety of other functional groups.
• explain the general methods of preparation of saturated mono carboxylic acids; • explain the industrial laboratory preparation of acetic acid • conduct an experiment to prepare acetic acid in the laboratory	• Preparation	Students should understand preparation of carboxylic acids. You may pick few appropriate chemical reactions to illustrate their preparations. 1. Oxidation of Primary alcohols $R\text{-}CH_2OH \xrightarrow{KMnO4} R\text{-}COOH$ 2. Oxidation of alkyl benzenes $Ar\text{-}R \longrightarrow Ar\text{-}COOH$ Oxidation is the most direct is generally used when possible, some lower aliphatic acids being made from the available alcohols, substituted aromatic acids from substituted toluenes.
• name write structural formulas of some fatty acids	• Fatty acids	Students should know fatty acids are carboxylic acids with a long hydrocarbon chain that are found in all cells. The hydrocarbon chains of animal fatty acids are more saturated than those of vegetable origin. With only a few exceptions, the fatty acids are all straight-chain compounds. Most fatty acids contain an even number of carbon atoms.
• describe some uses of common carboxylic acids	• Uses	Let the students mention some uses of common carboxylic acids like acetic acid
• list common sources of esters;	**6.2 Esters** (6 periods) • Sources of esters	Students should appreciate that esters are found abundantly in living organisms.
• write the general structural formula of esters	• Nomenclature	Students should know that an ester is made by replacing the H of an –OH acid

Competencies	***Contents***	***Suggested Activities***
• write the molecular formulas names of some simple esters		with an R group. Carboxylic esters derived from carboxylic acids, therefore, have the formula RCOOR'. They are named as derivatives of the corresponding carboxylic acids, the best way to begin is by 1. Naming the acid, using either the IUPAC name or the common name. 2. After that, the procedure is the same as for naming salts. The "-ic acid" is changed to "ate," this word is the second word in the ester name. 3. The first word is the name of the R' group. E.g.  $CH_3\text{-}\underset{\underset{O}{\\|}}{C}\text{-}O\text{-}CH_2CH_3$ 1. Name of acid: Ethanoic acid or acetic acid 2. Change "-ic acid to "ate,": Ethanoate or Acetate 3. Name of R' comes first: Ethyl Hence full name is: Ethyl ethanoate or Ethyl acetate. Ask students to give names for few common carboxylic esters. (Remember there are inorganic esters like phosphoesters as well.) You may ask them also to identify the acid alcohol (or phenol) portions of these carboxylic esters.
• describe physical properties of esters;	• Physical properties	**Boiling points**: Ask students to guess whether esters will have higher, lower or similar boiling points compared to compounds of comparable molecular mass that do form hydrogen bonds. Because ester molecules cannot form hydrogen bonds with each other, esters have lower boiling points than compounds of comparable molecular mass that are capable of forming hydrogen bonds.
• explain the chemical properties of esters;	• Chemical properties	**Solubility**: Since carboxylic esters can form hydrogen bonds with water, it is not surprising that their solubility in water is about the same as that of carboxylic acids of the same molecular weight.

Competencies	*Contents*	*Suggested Activities*
• explain the general methods of preparation of esters • describe some uses of common esters;	• Preparation • Uses	**Odor**: In sharp contrast to the disagreeable odors of carboxylic acids, esters have pleasant odors. Many of the odors of fruits flowers result from mixtures of carboxylic esters, many of them are used in perfumes food flavorings. Students should be aware that the most important reaction of carboxylic esters is their hydrolysis, which may be catalyzed by either acids or bases. For synthetic purposes, base catalysis is often preferred because the reaction is not reversible. Base catalyzed ester hydrolysis is called saponification. In biological systems, ester hydrolysis is catalyzed by enzymes. Students should know preparation of esters. Carboxylic esters can be considered as a combination of a carboxylic acid an alcohol (or phenol). In fact, they can be synthesized by heating a mixture of a carboxylic acid an alcohol in the presence of an acid catalyst such as H_2SO_4. Students should appreciate that esters have numerous uses: - Solvent - Medicine - Clothing eg polyesters (Dacron) - Fragrance in perfumes - Plasticizers eg octyl phthalate
• define fats oils • write the general structural formula for fats oils • write the structures of some common triglycerides	**6.3 Fats Oils** (6 periods) • Structure	Fats oils are esters. Students should be able to see that they are widely found in nature especially in living things as esters of glycerol, in fact triacylglycerols. Fats oils are triacylglycerols whose hydrocarbon tails, derived from carboxylic acids, are long enough to make them hydrophobic. As the hydrocarbon tail gets longer, they become more fatty i.e. less water soluble. The degree of unsaturation of the fatty acid tail portion of triacylglycerol is responsible for the physical states of the fats oils.

Competencies	***Contents***	***Suggested Activities***
• describe physical properties of fats oils • explain hardening of oils (process of converting oils to hard fats)	• Physical properties • Hardening of oils	Students should know that animal fats are generally solids at room temperature those from plants are usually liquids (oils). Warn students not to confuse the edible vegetable oils with mineral oils which are entirely different; mineral oils (petroleum) which are long chains of alkanes. Giving local examples would bring chemistry closer to the students this should be done when situations allow. It would be interesting to show how hydrogenation of vegetable oils results in the formation of solid fat (eg SHENO LEGA) which is not different in composition from animal fat. Ask students if it is OK to consume Sheno Lega during fasting season. It is also useful to show how to calculate oxidation numbers of the various carbon atoms in both fats oils use oxidation number as an indicator of energy content; the lower the oxidation number the higher the energy content of the substance. Reference could also be made to the type of hybridization of the carbon atoms. Vegetable oils differ from animal fats in possessing higher degree of unsaturation, but could have double bonds. You should ask the students to check if they know that numerous fats contain sites of unsaturation like vegetable oils. This is a common misconception in students at this level. It is known, other things being equal, unsaturation in fat tends to lower its melting point thus tends to make it a liquid at room temperature. Hydrogenation of some of the double bonds in oils converts these liquids into solids having a consistency comparable to that of butter. The accompanying change of state is called hardening of oils.
• explain rancidity	• Rancidity	Students should be aware that hydrogenation not only changes the

Competencies	*Contents*	*Suggested Activities*
		physical properties of a fat, but also- this even is more important-changes the chemical properties: a hydrogenated fat becomes rancid much less readily than does a non-hydrogenated fat. Rancidity is due to the presence of volatile, bad-smelling acids aldehydes. Here you could engage your students with discussion in groups on concepts related to reactivity of alkenes vs alkanes. Fats oils become rancid if left un refrigerated. This can be used to show how the numerous bacteria produce enough catabolic enzymes that break the long fatty acids into smaller pieces of lower boiling point that vaporize at room temperature, hence the bad odor. Mothers purify butter (animal fat) by heating it to boiling which kills the bacteria also denature the enzymes responsible for rancidity of butter.
• define soap	• Soaps detergents - Soap	Students should be aware that the making of soap is one of the oldest of chemical syntheses. Soaps are the alkali salts of fatty acids. (It is not nearly as old as the production of ethyl alcohol; may be his desire to get intoxicated is older than his desire to be clean.) Soap may vary in composition method of processing: - alcohol can be added to make it transparent; - air can be beaten in to it to make it float; - perfumes, dyes, germicides can be added; if potassium salt (instead of sodium salt), it is soft soap Chemically, however, soap remains pretty much the same, does its job in the same way.
• define detergent	- Detergents	Of the straight-chain primary alcohols obtained from fats-or in other ways-the C_{12} to C_{18} alcohols are used in enormous quantities in the manufacture of detergents (synthetic soap- like molecules). Although the synthetic

Competencies	***Contents***	***Suggested Activities***
		detergents vary considerably in their chemical structure, the molecules of all of them have one common feature, a feature they share with ordinary soap: they are amphipathic, have a large non-polar hydrocarbon end that is oil-soluble, a polar end that is water-soluble. The C_{12} to C_{18} alcohols are converted into the salts of alkyl hydrogen sulfates when treated with H_2SO_4. The alkyl sulfates, when treated with NaOH, produce the detergent, sodium alkyl sulfate. There are also non ionic detergents to be used when the ionic detergents are not effective. These detergents act in essentially the same way as soap does. They are used because they have certain advantages. For example, the sulfates retain their efficiency in hard water, since the corresponding calcium magnesium salts are soluble. Ask the students if they know what makes water temporary or permanent hard water. (This can be raised again in grade 12 chemistry in the context of hydrolysis of neutral salts basic salts: detergents are neutral whereas soaps are basic.) But detergents are not without drawbacks: some are non biodegradable, which makes them environmentally non-friendly. (The new detergents are better in being more biodegradable than the older generation.)
• explain saponification • prepare soap • explain the cleaning action of soaps	- saponificaiton (soap-making) - Cleaning action of soap	Students should be aware that soap making (saponification) is one of the oldest chemical technologies. Man first boiled goat tallow wood ash to give a lathering cleansing product. The goat tallow contained the ester, the wood ash the necessary alkali. Chemicals with same reactive groups are the main raw materials still used in traditional modern soap manufacture. Students should know that saponification is the chemical process of an alkali breaking down an alkyl alkanoate or ester to form soap. Demonstrate preparation of soap. The students should also appreciate the main continuous processes for manufacturing

Competencies	*Contents*	*Suggested Activities*
		of soap in modern industries. Students should be able to see the differences similarities between Soaps Detergents. They should also be guided to be aware of the advantages disadvantages of soaps detergents in relation to hard water biodegradability of detergents. At this point ask students if they expect soaps detergents to be water soluble. If they answer in the affirmative help them see that it is the polar head that associates with polar water-hydrophilic. The non-polar end is water-insoluble, is thus hydrophobic (or lipophilic). Such molecules are called amphipathic: they have both polar non-polar ends , in addition, are big enough for each end to display its own solubility behavior. In line with the rule of “like dissolves like”, each non-polar end seeks a non-polar environment; in this situation, the only such environment about is the non-polar ends of other soap molecules, which therefore clump together in the center of the micelle. Negatively charged carboxylate groups bind to the surface of the micelle it is surrounded by an ionic atmosphere. Water alone cannot dissolve fat grease that make up contain the dirt; oil droplets in contact with water tend to coalesce so that there is a water layer oil layer. But the presence of soap changes this. The non-polar ends of soap molecules dissolve in the oil droplet, leaving the carboxylate ends projecting into the surrounding water layer. As can be seen, this emulsifying, hence cleansing, property is not limited to carboxylate salts, but is posed by other amphipathic molecules.

Assessment

The teacher should assess each student's work continuously over the whole unit 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: List common organic acids and mention their sources; write the general formula of saturated monocarbocylic acids, the molecular formulas and names of the first six members and the structural formulas for the first four members; give examples for monocarboxylic, dicarboxylic and tricarboxylic acids; explain the general methods of preparation of acetic acid, and conduct experiment to prepare acetic acid in the laboratory; describe the physical and chemical properties of saturated monocarbaxylic acids; name and write structural formulas of some fatty acids; describe some uses of common carboxylic acids; list common sources of esters; write the general structural formula of esters and the molecular formulas and names of some simple esters; explain the general methods of preparation of esters; describe the physical and chemical properties of esters, and some uses of common esters; define fats, oils, soap and detergent; write the general structural formulas of fats and oils, and the structures of some common triglycerides; describe physical properties of fats and oils; explain rancidity and hardening of oils, explain saponification and prepare soap, and explain the cleaning action of soap.

Students above minimum requirement level

Students working above the minimum requirement level should be praised their achievements recognized. They should be encouraged to continue working hard 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 additional lesson time during breaks or at the end of the day.