CHEMISTRY

TEACHER'S GUIDE

GRADE 9

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Introduction

The study of chemistry at this cycle, Grades 9 -10, prepares your students for the future nation building, both practically and philosophically. 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 efficiently and effectively, your students come to understand how chemistry deals with the daily and routine lives of theirs and the citizens at large. On a higher level, the role of chemistry is also significant nationally as well as internationally.

To materialize the above stated major goals, encourage the 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 range scope of the development activities of their nation.

At this cycle, the students are 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. To this end, the specific objectives of chemistry learning at this cycle are to enable them to:

- gain a solid knowledge of chemistry.
- appreciate the power, elegance and structure of chemistry.
- use chemistry in their daily life.
- understand the essential contributions of chemistry to the fields of engineering, science, 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 The traditional teaching-learning paradigm has been replaced by active and participatory student-centered model. A student-centered classroom atmosphere and approach stimulates student's inquiry. Your role in such student-oriented approach would be a mentor who guides the student construct their own knowledge and skills. A primary goal when you teach a concept 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.

You are also encouraged to motivate students to develop personal qualities that will help them in real life. For example, student-oriented teachers 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 they could motivate addressing their material to groups and to present themselves and their ideas well. Support students and give them chance to stand before the class and present their work.

Similarly, help them by creating favorable conditions for students to come together in groups and exchange ideas about what they have learned and about material they have read. In this process, the students are given many 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 your teaching time as you plan how to approach a topic. The guide suggests teaching-time periods for each subject you will teach. The guide also contains answers to the review questions at the end of each topic.

Each section of your teacher's guide includes student-assessment guidelines. Use them to evaluate your students' work. Based on your conclusions, you will give special attention to students who are working either above or below the standard level of achievement. Check each student's performance against the learning competencies presented by 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, give them extra help. For example, you can give them supplementary presentations and reviews of the materials in the class. Extra time to study, and develop extra 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. Search for new web sites is well worth your time as you investigate your subject matter. 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 them into practice in your classroom. Use your own full capacity, knowledge and insights in the same way as you encourage your students to use theirs.

General Information to the Teacher

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

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

1. Organizing groups

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

2. Discussion

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

- follow up how every student participates in the discussion.
- be part of the discussion in some groups for a few minutes and see how the discussion among students is going on.
- give assistance and guidance when students are in need.
- give them hints when they face difficulties or have questions on the points suggested in the activities.

- ask questions related to the points in the activity to facilitate the interaction among students during discussion.

3. Presentation

Students are expected to present:

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

Rationale for active learning

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

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

A. Gapped lectures

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

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 atoms.

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. Demonstration

This is a method where the teacher shows the students how something is done. For example, preparation of iron sulphide from sulphur and iron.

E. 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 unit 3 – 5 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 of matter into subatomic particles in unit one.

G. Question and answer (inquiry)

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

H. Spider diagram

Students write a topic at the middle and write ideas related to the topic around the topic and draw a line connecting each idea to the central idea. For example, factors affecting rate of reactions.

I. 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 chemical bonding, periodic table.

J. 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 importance of periodic classification.

K. Drawing a picture, map or graph

This is a very useful way for visual learners to internalize, knowledge, concepts and information. For example, this method can be applied in unit 2, 5 etc.

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.

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Unit 1 Structure of the Atom

Unit Overview **Total Period Allotted 15**

Unit One has five parts. The first part, **Section 1.1**, introduces the concept of atomic theory. It also presents Dalton's atomic theory and the modern atomic theory.

Section 1.2 describes the discoveries of the fundamental subatomic particles and the nucleus. It presents the discovery of the electron, properties of cathode rays, the charge and mass of the electron, radioactivity, and the discovery of the atomic nucleus. In addition, this section deals with Rutherford's contribution to the structure of an atom (along with his associates, Hans Geiger and Ernest Marsden), and the discovery of the neutron by James Chadwick.

Section 1.3 presents the composition of atoms and isotopes. It includes detailed information about atomic number, mass number, isotopes, atomic mass, and the designation of isotopes, relative atomic mass, and average atomic mass of elements.

Section 1.4 presents atomic models. It covers Dalton atomic model, Thomson model, Rutherford model, Bohr model and the quantum mechanical model. The last part of this section deals with the distribution of electrons among available sub-shells in an atom.

Please encourage your students to interact and give them time to grasp the topics. Also, be sure that they give full attention to the introductory activity in each section. This activity prepares them for the topic they are about to study. All sections include examples and activities that give your students many chances to practice what they learn.

The methodologies to be implemented for teaching the contents in this unit are gapped lecture, answer and question, group discussion, visual-based active learning, drawing a map (picture) etc.

Unit Outcomes

After completing this unit, students will be able to:

- comprehend Dalton's atomic theory and the modern atomic theory;

- understand the discovery of the electron and the nucleus;
- know terms like atomic number, mass number, atomic mass, isotope, energy level, valence electrons and electron configuration;
- understand the Dalton, Thomson, Rutherford, Bohr and quantum mechanical atomic models;
- develop skills in
    * determining the number of protons, electrons and neutrons of atoms from atomic numbers and mass numbers,
    * calculating the atomic masses of elements that have isotopes,
    * writing the ground-state electron configurations of atoms, using sub energy levels and drawing diagrams of atoms.
- demonstrate scientific inquiry skills: observing, comparing and contrasting, communicating, asking questions, and applying concepts.

Main Contents

1.1 Atomic theory.

1.2 Discoveries of the fundamental subatomic particles and the atomic nucleus.

1.3 Compositions of an atom and the isotopes.

1.4 The atomic models.

Answers to Review Exercises

1.1 Atomic Theory

Period Allotted 3

Competencies

After completing this section, students should be able to:

- describe Dalton's atomic theory;
- describe the modern atomic theory;
- compare and contrast Dalton's atomic theory and the modern atomic theory.

Forward Planning

Read the contents on atomic theory from the students' text and make the necessary preparation. You also need to read the contents in this guide to get clear information about the suggested activities. Plan how to manage students when they discuss the start-up activity and Activity 1.1 – 1.3. Have a detailed time budget plan for all activities you perform during each period. Have a list of students in each section so that you can use it to record the performance of every student in the section.

Subject Matter Presentation

The methodologies to be implemented for teaching the contents in this topic are group discussion, gapped lecture and inquiry.

You are advised to begin the unit with the start-up activity. The start-up activity is designed to illustrate the atomic model. First, let the students discuss the activity in groups for a few minutes. Then choose a few groups to present their ideas to the class. After the students' presentation, harmonize the discussion with your conclusion.

Tell your students that:

- The layers peeled off the onion represent energy levels in the atomic structure where the electrons reside.
- The core represents the nucleus.

Dalton's Atomic Theory

You are advised to use group discussion and gapped lecture for the lesson. Start the lesson by asking the students to explain what an atom is. Next, introduce the ways that Democritus and Aristotle conceived of the atom.

Begin teaching the contents in the lesson with activity 1.1 (given in the student textbook). Let the students participate and discuss in groups. Then invite some students to present their conclusions to the class. After their presentations, hold a whole-class discussion. After the discussion, please harmonize the concepts suggested by students with the facts. The students should clearly see that the postulates of Dalton were based on experimental evidence whereas those of the ancient Greeks were simply a result of thinking and reasoning.

Let students know that Dalton's atomic theory was developed in 1803. Emphasize that it is considered the basis of modern physical science.

Inform the students that Dalton proposed the atomic theory based on the ideas of elements, compounds and the three laws of chemical combination: the law of conservation of mass, the law of definite proportion and the law of multiple proportions. Let the students be familiar with the concepts of the three laws. Then continue introducing Dalton's atomic theory.

After introducing Dalton's atomic theory, ask the students to guess which of its postulates were discovered to be wrong and to justify their opinions. Following their responses, introduce the discredited postulates and explain why they are wrong.

Activity 1.2 is designed to help students to differentiate between the law of definite proportions and the law of multiple proportions. Students should be able to understand the two laws based on concrete examples. Let them discuss the activity in groups for some minutes and then present their opinions to the class. After presentation, harmonize concepts as follows:

For example, sodium chloride, NaCl is composed of the elements sodium and chlorine only. These elements are always in the proportion of 39.34 % sodium to 60.66 % chlorine. This verifies the law of definite proportions.

For the law of multiple proportions, water and hydrogen peroxide can be considered. Assume that each of these compounds contain 1.0 g of hydrogen. In H_2O, 8 g of oxygen combines with 1 g of hydrogen. In H_2O_2, 16 g of oxygen combine with 1 g of hydrogen. The ratio of the masses of oxygen in H_2O and H_2O_2 is exactly 1:2 respectively. This verifies the law of multiple proportions.

After completing the presentation of Dalton's atomic theory, proceed to Activity 1.3. This activity is designed to assist students discover the relationship between Dalton's atomic theory and the three laws of chemical combinations. So let the students discuss Activity 1.3 in groups and then present their views to the class. Then, harmonize concepts.

Dalton's theory is useful for explaining the 3 laws of chemical reactions.

i. **The law of conservation of mass**:- Since a chemical reaction involves only the separation, union, and rearrangement of atom, mass is conserved in a chemical reaction.

ii. **The law of definite proportions**:- According to Dalton's theory, since atoms combine in small whole number ratios, a compound should contain the same elements combined in a definite ratio by mass of the elements.

iii. **The law of multiple proportions**:- Since atoms combine in whole number ratios, the masses (number of atoms) of one element combining with a fixed mass (number of atoms) of the other elements must be in whole number ratio. A pair of compounds like (H_2O and H_2O_2), (CO and CO_2) illustrate this law.

At the end of this section, introduce the modern atomic theory. Let the students suggest the postulates of modern atomic theory by correcting the wrong ones from Dalton's atomic theory.

Assessment

Assess each student's work throughout the section. Using students' performance list, record how every student takes part in discussing the suggested Activities 1.1 – 1.3, presentation after discussion, answering questions during harmonizing concepts or stabilizations. You can also give Exercise 1.1 as class work or homework. Check their work and record how every student answered the questions accordingly. You may also prepare questions of your own which are relevant to this content for assessment. From your records, check whether or not the suggested competencies are achieved.

Additional Questions

* 1. Why were Democritus' ideas rejected by other philosophers of his time?

2. Explain how Dalton's atomic theory explains the fact that mass is conserved in chemical reactions.

Answers to Additional Questions

1. Democritus believed that atoms could not be created, destroyed, or further divided. However, after a certain period of time, Democritus ideas met criticism from other philosophers. Some of their challenges were 'what holds the atoms together?' and 'how do atoms move through space?'. Since Democritus was unable to answer the philosophers' questions, his atomic theory was rejected eventually.
2. Dalton studied numerous chemical reactions, and was able to accurately determine the mass ratios of the element involved in the reactions. Based on the experimental observations, Dalton's atomic theory easily explains the conservation of mass in chemical reactions as the result of the separation, combination, or rearrangements of atoms. Because the atoms are rearranged in the chemical reactions, their masses are conserved.

Answer to Exercise 1.1

1. The basic postulates of Dalton's Atomic Theory are:

i All elements are made up of small particles called atoms.

ii Atoms are indivisible and indestructible.

iii All atoms of a given element are identical in mass and in all other properties.

iv Atoms are neither created nor destroyed in chemical reactions.

v Compounds are formed when atoms of more than one element combine.

vi In a given compound, the relative numbers and types of atoms are constant

2. A comparison of Dalton's theory and the modern atomic theory

Dalton's atomic theory	Modern atomic theory
1. Elements are made up of small particles called atoms.	Elements are made up of small particles called atoms.
2. Atoms can neither be created nor destroyed.	Atoms cannot be created or destroyed during ordinary chemical reactions.
3. Atoms are indivisible	Atoms are divisible. They can be subdivided into electrons, protons and neutrons.
4. All atoms of the same element are identical and have the same mass and size.	All atoms of the same element have the same atomic number but may vary in mass number due to the presence of different isotopes.
5. Atoms of different elements have different mass and size.	Atoms of different elements are different.
6. Atoms combine in small whole numbers to form compounds.	Atoms combine in small whole numbers to form compounds.

The similarities and differences of the two theories are presented in the table given above. Please study them carefully and help your students to understand the ways in which postulates number 2, 3, and 4 were modified.

1.2 Discoveries of the Fundamental Subatomic Particles and the Atomic Nucleus

Period Allotted 3

Competencies

After completing this section, the students will be able to:

- explain the discovery of the electron,;
- explain the discovery of the nucleus;
- explain the discovery of the neutron.

Forward Planning

You are expected to read the contents on discoveries of the fundamental subatomic particles and the nucleus from the students' text or any other reference materials to make appropriate preparation. Read the contents in this guide to help you how to prepare yourself for the teaching-learning process for the section. Prepare a detailed plan that shows the contents and activities you deal with during each period and when to give students the project work suggested in this section. You also need to plan how to manage students when they discuss Activities 1.4 – 1.8. Prepare the diagrams of simple cathode ray tube, effect of cathode rays on the paddle wheel and Rutherford's experiment. Also plan how to budget your time for students to discuss the activities, for harmonizing concepts and for other activities you perform during each period.

Teaching Aids

Diagrams showing a simple cathode ray tube, the effect of cathode rays on a paddle wheel and the Rutherford's experiment.

Subject Matter Presentation

Discovery of the electron

You are advised to use group discussion, question and answer, visual-based active learning and collaborative learning as your teaching methods for this subunit. Use these methods in a way that promotes active learning. Encourage your students to participate in the teaching-learning process. Begin the lesson with Activity 1.4. Activity 1.4 is designed to give students an understanding of the composition of the atom, in terms of the subatomic particles, their location in the atom and their charges. Have the students discuss this activity for a few minutes in groups. Then invite some of the groups to present their ideas to the class. After presentation, harmonize concepts suggested by the students with the truth.

1. An atom consists of a very dense central nucleus containing positively charged protons and uncharged neutrons; and is surrounded by negatively charged electrons.
2. Emphasis should also be given to the fact that the discovery of the subatomic particles of the atom disproved Dalton's postulate of the indivisibility of the atom.

Then continue by using gapped lecture to introduce the discovery of the electron. Introduce the scientists who attempted to discover the electron and describe the apparatus they used, their observations and conclusions. Give the students an activity related to the points you explained and check how they are doing.

Then, ask students if they know anything that works in the same way as a discharge tube. After they answer the question, tell them that fluorescent lamps and television tubes are good examples. Continue by explaining properties of cathode rays. Ask the students if they know why cathode rays have such properties.

After you finish presenting the properties of cathode rays, let the students construct the suggested model and then have each group show the model to the class. Then, continue your discussion with the ideas of the charge and mass of an electron. Use Activity 1.5 to start introducing the concepts. The activity enables students to realize how the determination of the e/m ratio confirmed the presence of subatomic particles. So, let the students discuss activity 1.5 in groups for a few minutes and one or two groups present their opinions to the class. After the presentations, harmonize concepts as:

In Thomson's experiment, it is observed that the charge-to-mass ratio of the electron is the same, regardless of the type of the gas in the tube. This led to the conclusion that the electron is a fundamental subatomic particle present in all matter with a characteristic charge and mass. Let students know how the charge-to-mass ratio and the charge of an electron were determined and introduce their magnitudes.

Discovery of the Atomic Nucleus

To teach the contents in this topic, use group discussion and gapped lecture as method of teaching. After referring the main points treated in the previous lesson, proceed by presenting the discovery of the atomic nucleus.

Start the lesson by defining what radioactivity is. Tell them the scientists involved in discovering and understanding this phenomenon. Introduce the types of radiation emitted during radioactive decay (alpha, gamma and beta rays) and their properties. Then, continue with the suggested activity.

Activity 1.6 is designed to help the students understand the similarities and differences between the three fundamental subatomic particles and the three types of rays, α , **ß** and **γ** rays. So, let them discuss activity 1.6 in groups for a few minutes and some groups present their conclusions. After the presentation, harmonize concepts.

Three types of rays are produced by the decay or breakdown of radioactive substances such as uranium. They are alpha rays, beta rays and gamma rays.

- alpha (α) rays or α- particles consist of positively charged particles.
- beta (ß) rays or ß particles are electrons, with negative charge, and are deflected by the negatively charged plates.
- gamma (**γ**) rays are high energy rays. They have no charge and are not affected by an external field.

You can use the following tables to compare and contrast the nature (properties) of electrons, protons, neutrons with α , **ß** and **γ** rays.

Particles	Symbol	Charge	Relative mass
Electron	e^-	-1	0.0005486 ≈ 0
Proton	p^+	+1	1.007276 ≈ 1
Neutron	n^o	0	1.008665 ≈ 1

Types of emission	Symbol	Charge	Relative mass	Composition
Alpha particles	α	+2	4	Similar to the helium nucleus
Beta particles	ß	-1	1/1837	Identical to the electron
Gamma ray	**γ**	0	0	High energy radiation

When you introduce the discovery of the atomic nucleus, inform the students about the work of Ernest Rutherford, his observations and conclusions. To teach about the discovery of the nucleus, you can assign one student to act as Ernest Rutherford. Tell him to prepare a circular mirror at the center of which contains a very small region of black paint. Then let him place a black screen behind the mirror. Let him, hold a small light source (torch light) directed to the mirror. First, he should explain that the mirror represents a thin gold foil, the torch light a source of alpha particles and the black screen as a fluorescent screen. He can use that most of the rays from the torch light pass through the mirror as do alpha-particles through atoms. The light rays that hit the black paint on the mirror do not pass through it as do alpha-rays that strike the nucleus. Finally let him explain, Ernerst Ruthford’s conclusion about the atoms.

Appreciate the student for this attempt and continue harmonizing concepts.

After you complete introducing the discovery of the nucleus, continue with Activity 1.7. This activity is designed to help students realize that bombarding a metal with alpha particles other than gold will lead to the same conclusions as those of Rutherford. So, let the students discuss activity 1.7 for some minutes in groups and present to the class. After the presentation, harmonize concepts.

The same results are obtained if copper, instead of gold, is bombarded with alpha particles. Most of the alpha particles pass through the copper undeflected. Some are deflected by small angles. Only a few of the alpha particles that collide with the dense positive centre are deflected by a very large angle.

Finally, explain how the neutron was discovered by James Chadwick.

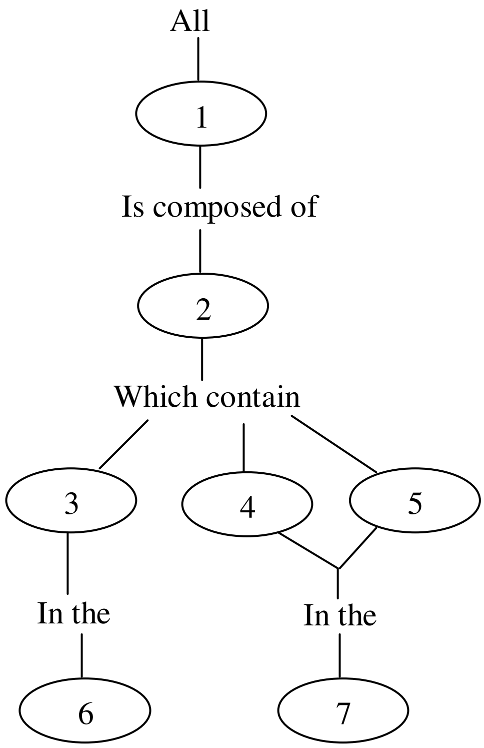
To conclude this section, let students do activity 1.8 in groups. Assist students to make a model of the hydrogen atom. Check their models and select the best one to use as a teaching aid.

Concept Mapping

After you introduce the structure of an atom, list the following terms on the blackboard and let some volunteer students try to complete the concept map using the terms. (Fill in the ovals with the following terms to create a concept map).

- electrons
- nucleus
- matter
- neutrons
- protons
- atoms
- empty spaces around nucleus

Concept mapping



Appreciate their attempts in completing the concept map and then harmonize using the appropriate terms as follows:

1. Matter
2. Atoms
3. Electrons
4. Protons
5. Neutrons
6. Empty spaces around nucleus
7. Nucleus

Assessment

You need to assess each student's work throughout section 1.2. Using students' performance list record how every student:

- participates in discussing activities 1.4 – 1.8
- takes part in presentations after discussions
- answers questions you ask during harmonizing concepts, stabilization or questions raised during discussions.
- answers questions given to them as class work or homework and quiz accordingly after checking their work (you can use exercise 1.3.

From your records, see whether or not the competencies suggested for the section are achieved by most of the students.

Appreciate students working above the minimum requirement level and encourage them to continue working hard. You can give them the additional questions suggested in this section. For students working below the minimum requirement level, give them extra attention in class or arrange extra lesson time. You may also set questions on contents in this section (1.2) in accordance with the suggested competencies and give them as an additional exercise so that they will catch up with the rest of the class.

Additional Questions

* 1. Describe the experimental basis for believing that the nucleus occupies a very small fraction of the volume of the atom.
* 2. Explain what keeps the electrons confined in the space surrounding the nucleus.

Answers to Additional Questions

1. Rutherford, with his associates, Geiger and Ernst Marsden, bombarded very thin foils of gold with α - particles, and they observed
   - The majority of α - particles penetrated the foil undeflected.
   - Some α - particles experienced slight deflections.
   - Only a few particles were either deflected by a very large angle or deflected back. This confirmed that the space (volume) occupied by the nucleus must be very small.
2. The electrons move through the empty space surrounding the nucleus. These electrons are held within the atom by the attraction with the positively charged nucleus. The volume of space through which the electrons move is extremely large as compared to the volume of the nucleus.

Answers to Exercises

Exercise 1.2

1. Two electrodes (cathode and anode) from a high-voltage source are sealed into a glass tube from which air has been evacuated. When the high voltage current is turned on, the glass tube emits a greenish light and a beam of light is seen at the anode. These rays flow from the cathode towards the anode in a straight line. They are called **cathode rays or electrons.**

2. J. J. Thomson determined the charge-to-mass ratio, ($e–/m$) of the electron, which finally helped to calculate the charge of the electron. The '$e–/m$' value was found to be 1.76×10^8 coulomb/g.

   Robert Millikan determined the charge of an electron ($e–$), using the oil-drop experiment. He found the charge of an electron to be 1.60×10^{-19} coulombs.

   Combinations of $e–/m$ and $e–$ values are used to determine the mass of an electron, which is found to be 9.11×10^{-31} kg.

Exercise 1.3

1. i) When an electric or magnetic field is applied in the path of cathode rays, they are deflected towards the positive plate in an electric field.

   ii) The properties of cathode rays show similar properties that are independent of the nature of the gas in the discharge tube and of the material used as the cathode.

2. Most of the space in an atom is empty.
3. The presence of a heavy positive centre in the atom, which Rutherford named the nucleus.
4. The charge of an electron.
5. a. C b. A c. B

1.3 Composition of the Atom and the Isotopes

Period Allotted 3

Competencies

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

- write the relative charges of an electron, a proton and a neutron;
- tell the absolute and relative masses of an electron, a proton and a neutron;
- tell the number of protons and electrons in an atom from the atomic number of the element;
- determine the number of neutrons from given values of atomic numbers and mass numbers;
- explain the terms isotope and atomic mass;
- calculate the atomic masses of elements that have isotopes.

Forward Planning

Make the appropriate preparation on composition of the atom and the isotopes by reading the contents in the students' text. You also need to read the teacher's guide on this section to plan how to present the lesson you intend to deal with during each period and get detailed information about the suggested Activities 1.9 – 1.15 and the methodologies you implement. Plan how to manage students during discussion, and also budget your time for every activity you perform. The time you allot for students to discuss activities, presentation after discussion, harmonizing concepts and other activities during each period should be indicated in detail.

Subject Matter Presentation

Composition of the Atom

We advise to use group discussion, question and answer methodologies for this topic.

After introducing the topic of the section, start the lesson using Activity 1.9. The activity is designed to help students realize that the atomic number (number f proton) of the element is always a whole number and the atomic number of every element is fixed, and it clearly distinguishes one element from another. Also the activity helps the students that atoms of the same element can have different masses and that atoms of different elements having same mass can exist. So, let the students discuss Activity 1.9 in groups for a few minutes. After they complete the discussion, invite some groups to present their conclusions to the rest of the class. After the presentations, harmonize concepts suggested by students with the truth.

1. Tell them the facts that, atoms consist of whole numbers of protons and electrons that are equal in number. No element could fit between Magnesium (atomic number 12) and Aluminum (atomic number13).
2. Help them understand that two atoms of calcium may differ in mass, due to different numbers of neutrons in their nuclei. Please inform the students of the compositions of the two nuclei as Calcium - 40 (20 p and 20 n) and Calcium - 44 (20 p and 24 n). Inform students that atoms of different elements may have the same mass because they have the same nuclear compositions for example Calcium - 40 (20 p, 20 n) and Potassium - 40 (19 p, and 21 n). However, strong emphasis should be given to the fact that Calcium - 40 and Potassium - 40 are different elements, because they differ in atomic number. Also, explain that atoms of different elements are different in atomic number and mass, as exemplified by Calcium - 40 and Cobalt - 59.

After harmonizing concepts, you can continue teaching the lesson by asking students to state the two regions in an atom. After their responses, tell them what the regions are.

To proceed further, ask them what the three fundamental particles of an atom are, where they are located in an atom, and their charges and masses. After appreciating the attempts of your students, give the appropriate answers to your questions. Next, let them do activity 1.10 in groups so that they can clearly see the regions in an atom and where the subatomic particles are located.

Assist students in sketching the model of the Chlorine atom. Check their models and select the best one as a teaching aid. After they complete this activity, continue with Activity 1.11. This activity helps students realize that atoms are extremely small, even though they are primarily composed of empty space. Therefore, let them discuss activity 1.11 in groups and present their findings to the class. Harmonize their group discussion with the following facts.

The empty space in an atom is at the microscopic level. A solid object consists of closely packed atoms, ions or molecules with very small spaces between them. These spaces are too small to pass hands through.

Atomic Number and Mass Number

To teach the contents in this topic, use group discussion, gapped lecture and question and answer as methods of teaching.

You can start the lesson using Activity 1.12. This activity is designed to enable students to differentiate neutral atoms, negatively charged ions and positively charged ions. Let them discuss the activity for a few minutes and then have some students from different groups present their findings to the class. After their presentations, harmonize concepts. First of all, remind the students the following important points:

- The numbers of electrons and protons in a neutral atom are equal.
- In a negatively charged ion, the number of electrons is greater than the number of protons.
- In a positively charged ion, the number of electrons is less than the number of protons.

i. A, B, E and G

ii. D and F

iii. Symbolic representation for B is ${}^{14}_{7}B$

Symbolic representation for D is ${}^{20}_{10}D^{-2}$

Symbolic representation for F is ${}^{11}_{5}F^{-}$

Next, ask the students to define atomic number and mass number. Following their responses, define the terms yourself. Tell the students how to designate an atom by a chemical symbol, atomic number and mass number. Let the students practice how to determine the number of protons, electrons and neutrons from the given values of atomic number and mass number.

Before you conclude the lesson, be sure that if students get the chance to discover why all atoms of a given element have the same atomic number. You can use Activity 1.13 for this purpose. So, let students discuss Activity 1.13 in groups for a few minutes and some groups present their conclusion to the class.

Then, harmonize concepts. Inform students that the atomic number of every element is fixed, and hence the chemical identity of an atom can be determined solely from its atomic number. "No *two elements can have the same atomic number*". For example, the atomic number 6 identifies the element carbon. No other element can have atomic number 6.

Isotopes and Atomic Mass

It is advisable if you use group discussion, question and answer methodologies.

You better start the lesson using activity 1.14. Activity 1.14 is designed to help students to have a clear understanding in determining the more abundant isotope among the isotopes of a given element. Let students discuss this activity in groups for a few minutes and then have some groups present their opinion. Following the presentations, harmonize concepts. Tell them that nitrogen, with a mass number 14, N-14, is more abundant because the atomic mass of nitrogen, 14.007, is nearer to 14.

Next, proceed to introduce isotopes and atomic mass. First, write the symbolic designation for isotopes of a given element with their mass numbers and atomic number on the blackboard. Allow the students to determine the number of protons, electrons and neutrons of the isotopes. Then let them define what isotopes are. Following their responses, define isotopes yourself and continue presenting relative atomic mass.

Let them differentiate actual atomic mass and relative atomic mass. Introduce them to the atom used as a standard to determine relative atomic mass. Give them the definition of atomic mass unit.

After that, let the students do Activity 1.15. Guide students when they construct the models of the three isotopes of magnesium. Check their models and select the best one to use as a teaching aid.

Finally, deal with average atomic mass of elements. Let students know what they should take into consideration when calculating average atomic mass. Solve some problems to acquaint students with the method for calculating average atomic mass. Give exercises and let them practice calculating average atomic mass.

Assessment

Assess each student's work throughout section 1.3. You can use students' performance list to record how every student

- participates in discussing Activities 1.9 – 1.15.
- takes part in presentations after discussions.
- answers questions you ask during mini-lecture, harmonizing concepts or stabilization.
- answers questions given as class work, homework and quiz accordingly after checking the work of every student. You can use Exercise 1.4, 1.5 and 1.6.

Check whether or not the suggested competencies for section 1.3 are achieved by the students. Appreciate students working above the minimum requirement level and encourage them to continue to work hard. Give extra attention to students working below the minimum requirement level in class or arrange extra lesson time. Give them questions recommended to them from the additional questions in this section or others of your own so that they will catch up with the rest of the class.

Additional Questions

1. Is the charge of a nucleus positive, negative or neutral? What about the charge of an atom?

* 2. Complete the following table.

Composition of different isotopes

Isotopes	Atomic number	Mass number	No of protons	No of neutrons	No of electrons
		32	16		
				24	20
Zn-64			30		
	9			10	
	11	23			

* 3. Calculate the atomic mass of magnesium, based on the information given in the table below.

Isotopic composition of magnesium

Relative masses (amu)	Percentage abundance (%)
23.985	78.99
24.986	10.00
25.982	11.01

Answers to Additional Questions

1. - The nucleus of an atom contains positive charge.
   - An atom is electrically neutral because its numbers of protons and electrons are equal.

2.

Composition of different isotopes					
Isotopes	Atomic number	Mass number	No of protons	No of neutrons	No of electrons
S-32	16	32	16	16	16
Ca-44	20	44	20	24	20
Zn-64	30	64	30	34	30
F-19	9	19	9	10	9
Na-23	11	23	11	12	11

3. The atomic mass of magnesium is calculated as:

$$23.985\times\frac{78.99}{100} + 24.986\times\frac{10.00}{100} + 25.982\times\frac{11.01}{100} = \underline{\underline{24.31}}$$

Answers to Exercises

Exercise 1.4

	Elements	Atomic number	Mass number	Number of neutrons
a	P	15	31	16
b	K	19	39	20
c	Fe	26	56	30

Exercise 1.5

1. The nuclear symbols for C-13 and C-12 are given as $^{13}_{6}C$ and $^{12}_{6}C$, respectively.
2. The numbers of protons, electrons and neutrons in the neutral atoms of **X**, **Y** and **Z** are given in the following table:

	Number of Protons	Number of Electrons	Number of Neutrons
$^{23}_{11}X$	11	11	23 – 11 = 12
$^{27}_{13}Y$	13	13	27 – 13 = 14
$^{31}_{15}Z$	15	15	31 – 15 = 16

Exercise 1.6

1. Isotopes are atoms of an element having the same number of protons but different number of neutrons. Isotopes of an element have identical chemical properties. The isotopes of an element can be represented by the symbol, ${}^{A}_{Z}X$. Isotopes of hydrogen have mass numbers of 1, 2 and 3, and are designated as ${}^{1}_{1}H$, ${}^{2}_{1}H$, and ${}^{3}_{1}H$ respectively.
2.

Isotopes	Protons	Electrons	Neutrons
${}^{234}_{92}U$	92	92	142
${}^{235}_{92}U$	92	92	143
${}^{238}_{92}U$	92	92	146

3. Protons = 82 , Electrons = 82, Neutrons = 124
4. Average atomic mass of thallium = 204.4
5. This is because the masses of the nucleus, even more than that of the electron, and so are small as compared to the radius of the atom. ($F = Km_1m_2/r^2$).
6. No. A nucleus with more than one proton and no neutron cannot exist. The positive charges of the protons repel each other. The neutrons reduce these electrostatic forces of repulsions.

1.4 Atomic Models

Period Allotted 6

Competencies

After the end of this section, the students should be able to:

- name the five atomic models;
- describe the Dalton, Thomson and Rutherford model;
- state Bohr's postulates;
- describe the Bohr model;
- describe the quantum mechanical model;
- describe main energy level and subenergy levels;
- define the term electronic configuration;
- write the ground-state electronic configuration of the elements;
- draw diagrams to show the electronic configurations of the first 18 elements;

- write the electronic configuration of the elements, using subenergy levels;
- write electronic configurations of the elements, using noble gas as a core; and
- describe valence electrons.

Forward Planning

Read the contents on the atomic models thoroughly from the students' text and other references materials, to know the concepts about Dalton's, Thomson's, Rutherford's, Bohr's and quantum mechanical model of the atom. Refer the teacher's guide on this section to plan which methodologies you can use and get information about the Activities 1.16 – 1.19. In addition, plan which content and activity you need to treat and the time allotment for every task you perform in detail during each period. Decide when to give them the task to prepare flash cards in groups that can be used when you teach them about electron configuration.

Teaching aids

- diagrams showing different atomic models
- flash cards

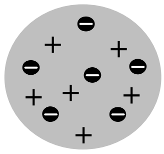
Subject Matter Presentation

You are advised to implement group discussion, gapped lecture as methodologies to teach this topic.

Start your presentation by reviewing the main points from the previous lesson. Let the students discuss Activity 1.16 in groups for a few minutes. After they complete, encourage some of the groups to present their views to the class. Activity 1.16 helps students to understand the historical background for the developments of the current atomic model. Ask them to think of different models of the atom as different models of cars. Ask them how car models differ from each other. Then have them consider how the different types of car models have changed through time. Students should gradually be informed about the developments of models of an atom as given in the student text, section 1.4

Proceed to teach the lesson by introducing the different atomic models. This should include Dalton's, Thomson's, Rutherford's, and Bohr's models. When you do so, ask students to suggest their views and tell what they know about each atomic model. First, introduce Dalton's model and then Thomson's model. After that, continue with Activity 1.17. The activity helps students to get a clear picture of Thomson's model. So, let the students discuss activity 1.17 in groups for a few minutes and then present their findings to the class. After the presentations, harmonize concepts using the following information.

1. The schematic drawing of a carbon atom using Thomson's model of the atom is given below. It is sometimes called the **plum-pudding model**.



2. Geiger and Marsden carried out an experiment in which the alpha rays collided against a thin metal foil. Their results showed that almost all incident alpha particles penetrated the foil and went straight, but a few were scattered at very large angles. The results of these alpha-particles scattering cannot be explained by Thomson's atomic model. This proved that Thomson's model of an atom was incorrect.

Next, continue to deal with Rutherford's and Bohr's atomic models. First, introduce to students how Rutherford proved that Thomson's model was incorrect. Then tell the students how he described an atom. After that, continue with Bohr's model. Emphasize Bohr's model and introduce the modifications he made to the existing concepts.

Let the students practice drawing diagrammatic representations of atoms. Let them practice determining the maximum number of electrons in a given main shell (main energy level) using the formula, ($\mathbf{2n^2}$). After introducing Rutherford's and Bohr's models, let the students discuss Activity 1.18 in groups and then present their opinion to the class. Following the presentations, harmonize concepts.

Rutherford's atomic model was the first to include a distinct boundary for positively and negatively charged particles of the atom. The Rutherford model of the atom has a small dense, positively charged nucleus around which electrons whirl at high speeds and relatively long distance from the nucleus where as Bohr proposed atomic model as electrons moving around the nucleus in circular orbits.

Bohr's atomic model accepted the quantum theory and successfully applied atomic theory to a hydrogen atom.

Finally, treat the quantum mechanical model. Explain how this model differs from Bohr's model. Introduce the sublevels (subshells) and the maximum number of electrons that each subshell can hold. At the end tell them to bring the flash cards they prepared for the lesson of the following period.

Electron Configuration

To teach this topic, use group discussion and gapped lecture as your methodologies.

You better start the lesson by defining what electron configuration is and what ground state electron configuration means. Then introduce the notation of quantum mechanical model and explain how it is used to describe the electron configuration of an atom. Give some examples so that students can understand how to use the notation. Activity 1.19 is designed to develop students' skill in writing electron configurations of an atom in the appropriate order, as given in the student's text book. So, let them do activity 1.19 in groups and present their findings to the class. Then, harmonize concepts as follows:

i. 2*p* subshell should be completely filled ($2p^6$) before the 3*s orbital.*
   Thus the correct electron configuration is $1s^2\,2s^2\,2p^6$
ii. 2*s* subshell should be completely filled ($2s^2$) before the 2*p* orbital
   Thus the correct electron configuration is $1s^2\,2s^2\,2p^5$
iii. The order of filing is wrong. 3*s* must come before 3*p*.
   Thus the correct electron configuration is $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$
iv. 3*p* orbital should be completely filled before the 3*d* and 4*s* orbitals.
   Thus the correct electron configuration is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2 3d^1$

Next, introduce the **Aufbau principle** or **building-up principle.** Let students practice writing electron configurations of elements using the diagonal rule. You better assist students to practice using flash cards.

Make sure that all groups have flash cards with names of subshells written on them. Student should write the names 1*s* on one flash card, 2*s* on another . . . up to 7*s* a total of about 16 cards. After that you can organize a competition among the groups. You can prepare many question of your own for the competition. The following are given as examples. (Students should show you only the flash card on which the appropriate answer is written).

1. In which subshell do electrons occupy after 3*p*?
2. What are the possible subshells available for a valence shell of an atom described by the principal quantum number, n = 4?
3. Which subshell has the lowest energy?
4. Which subshell is occupied by the last electrons of Calcium (Z = 20)?
5. Which subshells are occupied by electrons in Aluminium atom (Z = 13)?

You can use the following answers for the above sample questions.

1.	4*s*	4.	4*s*
2.	4*s*, 4*p*, 4*d*,4*f*	5.	1*s*, 2*s*, 2*p*, 3*s*, 3*p*
3.	1*s*		

Emphasize the quantum mechanical model and present the energy sublevels for the first four main energy levels. Before you conclude the lesson, give a brief summary of electron configurations, based on the rules of filling atomic orbitals. Also explain the method of writing electron configuration of a noble gas as a core.

Assessment

You need to assess how every student is working throughout section 1.4. Using students' performance list record how every student

- takes part in discussing Activities 1.16 – 1.19.
- involves in presenting concepts after discussion.
- answers questions raised during discussion, answers questions you ask during harmonizing concepts and stabilization.

Give them also Exercise 1.7 – 1.9 and the review exercise on unit 1 as class work or homework and a test. Check their work. See how many of them answered the questions accordingly and record their performances in your list.

Based on the record you made, check whether or not the suggested competencies for section 1.4 are achieved. Encourage students working above the minimum requirement level. Give recognition for their achievements and advise them to work hard. Let them also attempt the additional questions given in the teacher's guide in this section.

In case of students working below the minimum requirement level, give them more attention in class and also arrange an additional lesson time. Whenever it is important, you can give the questions recommended to them from the additional questions given in the guide. In addition to that prepare questions of your own based on the suggested competencies and give them as additional exercise and check their work.

Additional Questions

1. How was Rutherford's model of the atom an improvement compared to Thomson's model?

* 2. Which model of the atom is based on the law of planetary motion?

3 Why did scientists consider Rutherford's nuclear model of the atom incomplete?

4 What were the main defects of Bohr's model of the atom?

* 5. An atom has the electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$. If its mass number is 70, what will its number of protons and neutrons be?

* 6. Write the names of the elements that are represented by each of the following configurations.

   a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
   b. [Ar] $4s^1$
   c. [Ar] $4s^2 3d^6$
   d. [Kr] $5s^1 4d^{10}$
   e. [Xe] $6s^1$

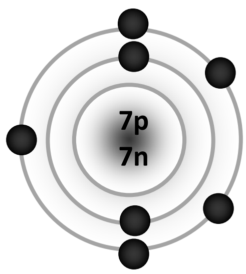
Answers to Additional Questions

1. Rutherford concluded that the Thomson plum pudding model was incorrect because it could not able to explain the results of the gold-foil experiment. Based on the experiment of the scattering α- particles, Rutherford confirmed that most of the mass and all of the positive charge of an atom are centered in a very small region called the nucleus.
2. Bohr's atomic model:- Bohr thought that the electrons orbit the nucleus in the same manner as planets orbit the sun.
3. Rutherford's model is said to be incomplete because his model lacked details on how the electrons of an atom are arranged in the space around the nucleus.
4. The following are the main defects of Bohr's theory:
   i. Bohr thought that electrons move in orbits around the nucleus. However today it is known that electrons do not actually move in orbits around the nucleus. Instead, they are described in terms of the probability of being found in certain regions of space around the nucleus. These regions of space are called orbitals.
   ii. His model was not successful for many-electron atoms.
5. Number of protons = 31; number of neutrons = 39
6. a. Bromine b. Potassium c. Iron d. Silver e. Cesium

Answers to Exercises

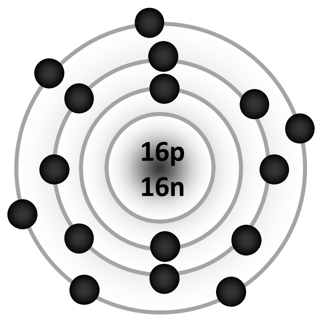
Exercise 1.7

A



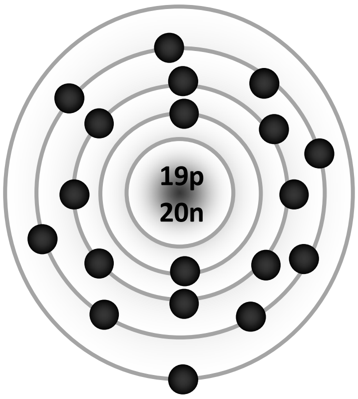
Nitrogen

b.



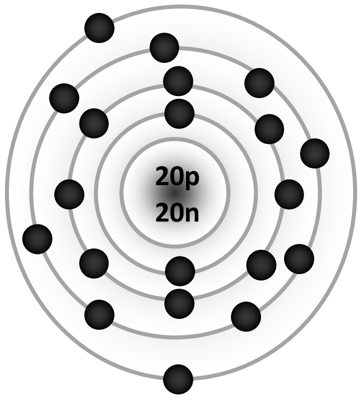
Sulphur

c.



Potassium

d.



Calcium

Exercise 1.8

Phosphorus ⇒ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Sulphur ⇒ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

Chlorine ⇒ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Argon ⇒ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Exercise 1.9

1. a. ${}^{16}_{8}O = [He]\ 2s^2 2p^4$

   b. ${}^{27}_{13}Al = [Ne]\ 3s^2 3p^1$

   c. ${}^{35}_{17}Cl = [Ne]\ 3s^2 3p^5$

   d. ${}^{40}_{20}Ca = [Ar]\ 4s^2$

2. a. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^2\ 3d^{10} 4p^6$

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

   c. $1s^2\ 2s^2\ 2p^6\ 3S^2\ 3p^6\ 4S^2\ 3d^{10}$

Answers to the Review Questions and Problems

Part I

1. a) True

   b) False

   c) True

   g) True

   h) False

d) False i) False

e) True j) True

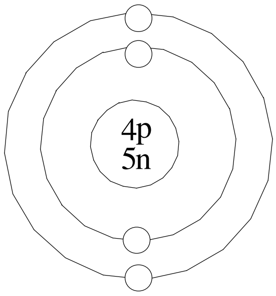
f) True k) False

Part II

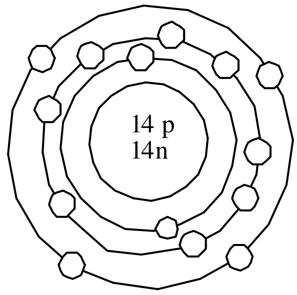
2. its nucleus 3. 10,000 4. $2n^2$

Part III

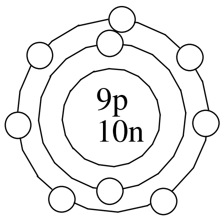
5. a)



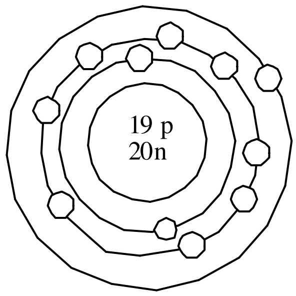
c)



b)



d)



6. The electron configuration is given as

a) $^{32}_{16}S$ = 2, 8, 6 or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ or [Ne] $3s^2\ 3p^4$

b) $^{52}_{24}Cr$ = [Ar] $4s^1\ 3d^5$

c) $^{56}_{26}Fe$ = [Ar] $4s^2\ 3d^6$

d) $^{64}_{29}Cu$ = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$ or [Ar] $4s^1\ 3d^{10}$

7. a) $3s^2\ 3p^4$, $4s^1$, $4s^2$ and $4s^1$ respectively

b) 16, 28, 30, 35 respectively

8. $3d^2$ is given in $_{22}Ti$ with [Ar] $4s^2\ 3d^2$

Therefore, a) [Ar] $4s^2\ 3d^2$

b) 2, 8, 8, 4

c) 22

d) 2

e) 22

9. 20.20

Part IV

10. The following modifications were made on the atomic theories of Dalton.

    i. Atoms are indivisible and indestructible.

    ii. All atoms of a given element are identical in mass and in all other properties.

11. Even though some modifications were made on Dalton's work, his theory was a brilliant and logical explanation of many experimental discoveries and laws that were known at that time. These scientific assumptions were very closely related to what is presently known about the atom. Due to this, John Dalton is often referred to as the father of modern atomic theory.

12. No, because the alpha-particles scattering experiment cannot be explained by Thomson's atomic model.

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Unit 2 Periodic Classification of the Elements

Unit Overview

Total Period Allotted 17

In this unit, the concept of periodic law and the periodic table are discussed based on early attempts to classify the elements. The modern periodic table, periodic law and the general trends of the periodic properties of elements in the periodic table are emphasized in the unit.

Section 2.1 deals with the historical development of the periodic table based atomic masses of elements.

Section 2.2 deals with the ideas of both Mendeleev and the modern periodic laws. Moreover, the classification of elements as representative, transition and rare earth elements, and also in terms of *s, p, d*, *f* - blocks of elements, are explained briefly.

Section 2.3 deals with the general trends in properties of the elements such as atomic size, ionization energy, electron affinity, electronegativity and metallic character across a period and down a group of the periodic table.

Section 2.4 describes the importance of the periodic table in the study of chemistry.

The methodologies which are recommended for teaching the contents of this unit are gapped lecture, group discussions, visual-based learning and answer and questions and drawing map (picture).

Unit Outcomes

After completing this unit, students will be able to:

- understand the periodic classification of the elements.
- develop the skills of correlating the electron configuration of elements with the periodicity of the elements, and in predicting the trends of periodic properties of elements in the periodic table.
- appreciate the importance of classification in chemistry.

- demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusions, applying concepts and problem solving.

2.1 Introduction

Periods allotted 1

Competencies

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

- *describe periodicity*

Forward Planning

Read thoroughly the contents about the introduction part from the students' text to help you make appropriate preparation. Read the contents in the teacher's guide as well to get clear information about the methodologies you implement to teach this section and the suggested start-up activity and Activity 2.1. Plan how to manage students during discussion and presentation, and make the teaching-learning process interesting. Set a plan to budget your time for group discussion, presentation, harmonizing concepts and other activities you perform during each period.

Subject Matter Presentation

To treat the contents of this section, it is advisable to use group discussion and inquiry as methods of teaching. After introducing the topic of the section, ask the students to suggest repeating events in their daily lives. Some examples are day and night, months of the year or days of the week. The daily-life examples will help the students to understand periodic events and periodicity.

Let the students discuss the start up activity for a few minutes. In the start-up activity, periodicity is shown to the students, using a calendar. This is a periodic table of days of the week. At the end of the start-up activity, students are expected to have a clear idea of the periodic table of the elements.

The students should be able to see that the calendar consists of families of days repeating after the seventh day and rows of weeks. An incomplete calendar is given to the students so that they can predict repeating events. What days are missing in a row and what name is given to columns of days?

Guide students in constructing the table and encourage them to predict the missing days and names of columns. When they complete their group presentation, inform them the missing days in columns 3(X) is 13, and columns 4(Y) is 14. They should be able to name columns 5 as Thursday.

You should also inform them that inclusion of X and Y completes columns 3 and 4. Inclusion of X and Y also completes row 3, which consists of days (11, 12, 13, 14, 15, 16, 17). In your further explanation, help them to see the previous month ended on Wednesday, and that the 28^{th} of that month was a Monday and that it would have appeared in the first row.

In your harmonizing session, please give students a similar understanding about the construction of the periodic table of the elements. Following the start–up activity, let the students discuss Activity 2.1 for a few minutes in groups and then have some groups present their opinions to the class. After the presentations, harmonize concepts.

In Activity 2.1**,** students should appreciate the fact that the modern periodic table is a result of many attempts. In this activity, group discussions and group presentations are directed to the historical development of the modern period table. During your explanation, let the students suggest what they already know about the early attempts of the classification of the elements. Inform them about the attempts made by Dobereiner and Newlands and explain how they made the classifications.

Reading Check: This is intended to encourage students to read other reference materials, analyze information, summarize and write reports.

Assessment

Assess how every student is doing throughout section 2.1. Record the performance of each student in the performance list you already have related to

- participation of each student in discussing the start-up activity and Activity 2.1.
- participation in presenting the ideas of the group after discussion
- answering questions you may ask during harmonizing concepts or stabilization

Give them Exercise 2.1 as class work or homework. Check their work and also record their performances. From the record, make sure that students have realized early attempts in classifying elements. Appreciate students working above the minimum requirement level and encourage them to continue working hard. You can give them the additional questions suggested in the guide. For students working below the minimum requirement level, you can set questions on Dobereiner’s triads and Newland’s law of octaves and give them additional exercise to assist them catch up with the rest of the class.

Additional Questions

1. Explain the contribution of Newland's law of octaves to the development of the modern periodic table.
2. What are the similarities between the Dobereiner's triads and Newland's law of octaves in the classification of elements?

Answers to Additional Questions

1. Because he proposed for the first time the periodicity in the properties of elements.
2. Both Dobereiner's and Newland's classified the elements on the basis of the atomic masses of elements.

Answer to Exercise 2.1

1. (a) First determine the average atomic mass of the middle element

$$\frac{9.0 + 87.6\ 1}{2} = 48.3$$

Since the actual atomic mass of calcium is 40, the elements Be, Ca, Sr do not follow the principle of Dobernier's triad.

(b) $\frac{6.94 + 39.01}{2} = 22.98$

Since this is very close to the atomic number of sodium, the elements: Li, Na, K, are in accordance with the concept of Doberner's triads.

2. During the time of Newlands, noble gases were not discovered. At present, the periodic similarity occurs in every ninth element due to the presence of noble gases.

2.2 The Modern Periodic Table

Periods Allotted-5

Competencies

After the end of this section, students will be to:

- state Mendeleev's periodic law;
- state modern periodic law;
- describe period;
- describe group;

- explain the relationship between the electronic configuration and the structure of the modern periodic table;
- describe the three classes of the elements in the modern periodic table;
- explain the four blocks of the elements on the basis of their electronic configuration in the modern Periodic Table;
- tell the block of an element from its electronic configuration;
- give group names to the main group elements;
- classify the periods into short, long and incomplete periods;
- tell the number of groups and periods in the modern periodic table;
- tell the number of elements in each period;
- predict the period and group of an element from its atomic number, and
- tell the block and group of an element from its electronic configuration.

Forward Planning

Read the contents on the modern periodic table thoroughly to familiarize yourself with the concepts in the section. Set a plan that shows the topics and activities you will treat during each period so that you can cover the entire content of the section within five periods. In addition read the teacher's guide to get information about the methodologies you need to implement and about the activities suggested in this section. In your plan, show the time allotted for every activity you perform during each period, such as group discussion, presentation, harmonizing concepts, gapped lectures etc. Make sure that your school has a periodic table chart. If not prepare the chart yourself.

Teaching Aids

- The Modern Periodic Table.

Subject Matter Presentation

The periodic law

It is preferable to implement group discussion, gapped lecture and visual-based active learning methodologies methods of teaching for this lesson.

You are advised to begin the lesson with Activity 2.2. Have the students perform and discuss it in their groups. Then invite some students to present their conclusions to the class. After their presentations, hold a whole-class discussion. After the discussion, please harmonize concepts suggested by the students with the truth as follows:

According to the given activity, there are 3 sets of horizontal boxes of elements that are filling the same valence shell. In the first set, 2 elements (H and He), in the second set, 8 elements (Li to Ne) and in the third set also 8 elements (Na to Ar) are grouped. The students should understand the regular patterns in forming sets of vertical and horizontal boxes. Elements in the same set of vertical boxes have the same outer electron configuration while those in horizontal boxes have the same number of shells.

Then, continue by teaching the concepts of Mendeleev's periodic table. First, ask the students to suggest the basis for the classification of elements in this table. After their responses, tell them that elements are classified based on their atomic masses. State Mendeleev's periodic law. Tell them why Mendeleev left some blank spaces in his table. Identify the elements for which Mendeleev left blank spaces. Explain the defects of Mendeleev's periodic table and continue by dealing with the modern periodic table. First, give students information about the contribution of Henry Mosely to the modern form of the periodic table. Be sure students are familiar with the basis for the classification of elements in the modern periodic table and the modern periodic law.

Characteristics of Groups and Periods

It is advisable to use gapped lecture as methodology in teaching this lesson. Start the lesson with questions by asking students to suggest their opinion. After getting responses, tell them the basis for the classification and define periods.

Explain the number of periods in the modern periodic table and also the number of elements classified in each period. After that, show them how they can determine the period number of an element from its electron configuration. Give them an activity and practice have them determining the period number of some elements. Check how well they are doing and give them corrections. Then, define groups and introduce the common features of elements in the same group. Show them how to determine the group number of an element form its electron configuration. To help students decide the position of an element in the modern periodic table, give them the following activity at the end.

The atomic numbers of four elements, W, X, Y and Z, are 11, 14, 8 and 16, respectively.

a) Determine the period and group number of each element.
b) Which elements are in the same period?
c) Which elements are belong to the same group?

When you are checking their work, make sure that the answers of the students are as follows.

a) W is in period – 3 and group I, X; in period – 3 and group IV; Y period – 2 and group VI and; Z in period – 3 and group VI
b) W, X and Z
c) Y and Z

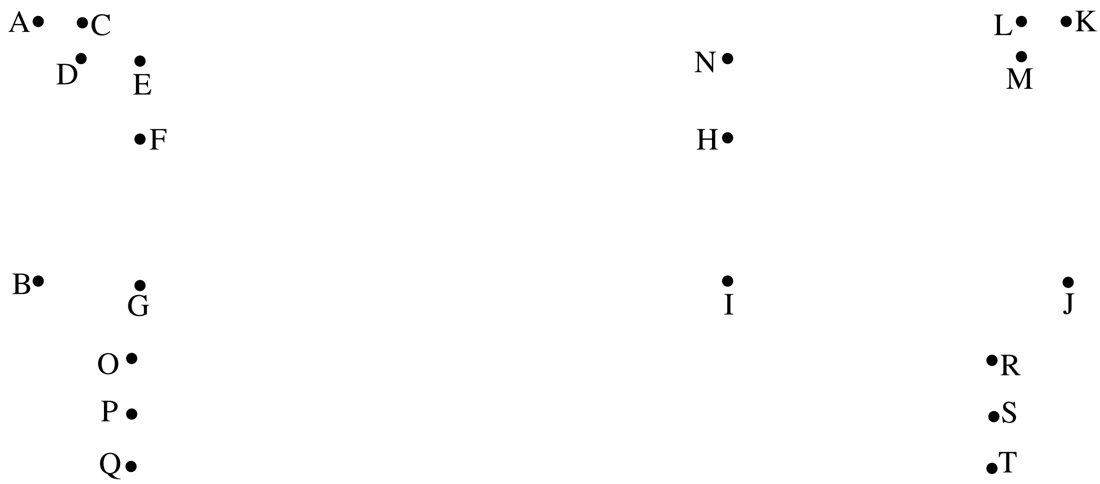
Classification of the elements

Use group discussion, visual-based learning and question and answer methodologies for this topic.

After reviewing the main points of the previous lesson, continue by presenting the classification of the elements. First, have the students perform the task given in Activity 2.3 in groups for a few minutes and then have some groups present their conclusions to the class. After the presentations, harmonize concepts. Use the following table for this purpose:

Element	Atomic number	Electron Configuration	Valence electron	Block
Nitrogen	7	[He] $2s^2\ 2p^3$	5	*p*
Sodium	11	[Ne] $3s^1$	1	*s*
Silicon	14	[Ne] $3s^2 3p^2$	4	*p*
Iron	26	[Ar] $3d^6 4s^2$	2	*d*
Zinc	30	[Ar] $3d^{10} 4s^2$	2	*d*
Krypton	36	[Ar] $3d^{10} 4s^2 4p^6$	8	*p*
Cerium	58	[Xe] $4f^1 5d^1 6s^2$	2	*f*

Then, proceed by introducing the classifications of elements as representative, transition and rare-earth elements. Give a gapped-lecture on the relationship between electron configuration and classification of the elements. To help students understand the concepts in your explanation, hang a big periodic table chart on the side of the black board. Let the students conduct group discussions about groups, periods, and the relationships between group and period number and electron configurations of elements, by using the chart as a teaching aid. Finally, check whether or not the students have realized the blocks of elements classified as representative, transition and rare-earth elements. Use the following information.



Without referring to the periodic table, let students join the points given by the letters and draw a modern periodic table. After that let them identify the regions that bound.

a. s - block elements	e. representative elements.
b. p - block elements	f. transition elements
c. d - block elements	g. f - block elements.
d. rare - earth elements.	

Harmonize concepts as follows:

a.	ACDEGB	e.	Refer a,b
b.	NIJKLM	f.	FGHI
c.	FGHI	g.	OQRT
d.	OQRT		

Reading check - This reading check is designed to help students to refer to additional materials, analyze information and write a short report.

Assessment

You are expected to assess each student's work throughout section 2.2. You can do this by using students' performance list and recording how every student:

- participates in discussing activity 2.2 – 2.3
- presents ideas after discussion
- answers questions raised during harmonizing concepts, stabilization and gapped lectures
- does the activity suggested in this guide related to classification of elements.

Give them also Exercise 2.2 and 2.3 as class work or homework, and quiz. Check their work and record their performances.

Based on the cumulative record, see whether or not the competencies suggested for this section are achieved by most of the students. Appreciate students working above the minimum requirement level. Give them the additional questions suggested for this section.

With regard to students working below the minimum requirement level, give them the necessary assistance in class and arrange additional lesson time when over required. You can also set questions of your own in accordance with the suggested competencies and give them additional exercise to help them catch up with the rest of the class. Use also question recommended to them from the additional questions.

Additional Questions

1. Luthar Meyer and Dimitri Mendeleev proposed similar periodic tables in 1869. However, Mendeleev is generally given more credit for the periodic table. Why? Explain.

* 2. Determine the group, period, and block of the elements with the electron configurations of:

a. [He] $2s^2 2p^1$

b. [Ne] $3s^2 3p^4$

c. [Kr] $5s^2$

d. [Ar] $4s^1 3d^{10}$

e. [Xe] $6s^1$

3. Explain why there are no *p* - block elements in the first period of the periodic table.

* 4. Write the electron configuration of the elements that corresponds to each of the following descriptions.

a) The group IIA element in the fourth period.
b) The noble gas in the fifth period.
c) The group VIA element in the second period.
d) The group IVA element in the third period.

Answers to Additional Questions

1. Mendeleev is generally given more credit than Meyer. This is because Mendeleev left blank spaces in his table for undiscovered elements and was able to predict the properties of these elements almost correctly.

2.

	Electron configurations	**Group**	**Period**	**Block**
a	[He] $2s^2 2p^1$	IIIA	2	*p*
b	[Ne]$3s^2 3p^4$	VIA	3	*p*
c	[Kr]$5s^2$	IIA	5	*s*
d	[Ar]$4s^1 3d^{10}$	IB	4	*d*
e	[Xe]$6s^1$	IA	6	*s*

3. The *p*-sublevel does not exist for the first principal energy level (n = 1).

4. a. The element is calcium, and its electron configuration is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
   b. The element is xenon (Xe) $\Rightarrow [Kr]5s^2 4d^{10} 5p^6$
   c. The element is oxygen $\Rightarrow 1s^2 2s^2 2p^4$
   d. The element is silicon $\Rightarrow 1s^2 2s^2 2p^6 3s^2 3p^2$

Answers to Exercises

Exercise 2.2

1. Noble gases (Group VIIIA elements) were missing in Mendeleev's periodic table because they were not discovered at the time of Mendeleev.
2. The first period corresponds to the first energy level (n = 1), and electrons are being added only to the 1*s* orbital of this energy level. Since the 1*s* orbital accommodates a maximum of 2 electrons, period 1 contains only 2 elements, with atomic numbers 1 and 2.
3. a) period 2 and group IVA
   b) period 2 and group VIIIA
   c) period 3 and group IIIA
   d) period 4 and group IA
   e) period 4 and group IIA
   f) period 3 and group VIA

Exercise 2.3

1. a) Na – $1s^2 2s^2 2p^6 3s^1$ - group IA ; period 3
   Ca – $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ - group IIA ; period 4
   Al– $1s^2 2s^2 2p^6 3s^2 3p^1$ - group III A; period 3
   b) Cl – $1s^2 2s^2 2p^6 3s^2 3p^5$ - group VIIA ;period 3
   S – $1s^2 2s^2 2p^6 3s^2 3p^4$ - group VIA ; period 3
   Ar– $1s^2 2s^2 2p^6 3s^2 3p^6$ - group VIIIA; period 3
2. $4s^2 4p^5$
3. a) Group IA, period 5 and *s*- block
   b) Group VIB, period 4 and *d*- block
   c) Group IVA , period 4 and *p*- block
4. **Critical thinking** - Hydrogen has only one valence electron, like all the rest alkali metals, and thus it is placed in Group IA.

2.3 Periodic Properties in the Periodic Table

Periods Allotted-6

Competencies

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

- explain the general trends in properties of the elements down a group of the periodic table,
- explain the general trends in properties of the elements across a period of the periodic table,
- deduce the properties of element from its position in the periodic table, and
- make a chart to show the trends in properties of the elements in the periodic table.

Forward Planning

Read all the contents on the periodic properties in the periodic table from the students' text and make the necessary preparation. Set a plan of your own that shows the contents and activities (2.4 – 2.15) which you treat during each period in such a way that you can cover the whole section within six periods. Your plan should also include the time allotted for every activity you perform during each period. Read the teacher's guide on this section to get information about the methodologies you implement for the section and to gain more ideas about each activity.

Collect data on melting points, boiling points, densities, and other properties from different reference materials, from your school library to provide your students with numerical values for use in their discussion. Plan how to manage students during discussions.

Teaching Aid

- Modern Periodic Table.

Subject Matter Presentation

Periodic Properties within a group

Apply group discussion and gapped lecture methodologies to teach this topic.

Please provide students with values of ionization energy, atomic size, electron affinity and electronegativity from the tables in the teacher's guide for Activity 2.4. Have them fill in these values and observe the trends in the properties of the elements across a period and down a group. After they fill the values, have the students discuss in groups.

Following their group discussion, have them present their findings and analyze any trend in the properties of the elements in a group and across a period. After the presentations, harmonize concepts. Tell them that they should come up with trends like the one shown below.

Down a group; ↓	**Across a period**; →
- atomic size increases	- atomic size decreases
- ionization energy decreases	- ionization energy increases
- electron affinity decreases	- electron affinity increases
- electronegativity decreases	- electronegaivity increases

After harmonizing concepts, proceed by introducing nuclear charge. First, ask the students to give their ideas about nuclear charge and effective nuclear charge. Following their responses, introduce them to the actual concepts of nuclear charge, effective nuclear charge and the shielding or screening effect.

Next, proceed by introducing Activity 2.5. This activity is designed to help students realize concepts related to nuclear charge, effective nuclear charge and the shielding effect. So, have the students perform the suggested task in Activity 2.5 in groups for a few minutes. Encourage one or two groups to present their findings to the class. After the presentations, harmonize concepts after drawing the Bohr's models for the following elements, as given in unit 1 of the student's text book.

Elements	**Shielding shells**	**Shielding electrons**	**Effective nuclear charge**
Beryllium	K-shell	2	+2
Magnesium	K and L- shells	10	+2
Calcium	K, L and M shells	18	+2
Lithium	K- shell	2	+1
Carbon	K- shell	2	+4
Fluorine	K- shell	2	+7

Finally, explain the trend in nuclear charge and effective nuclear charge in a given group. After introducing concepts of nuclear charge, continue by teaching atomic size or atomic radius. Before you present the details, ask them to define atomic size. Following their responses, give them the appropriate definition, supported by a diagram.

Next have them perform the suggested task in Activity 2.6 in groups for a few minutes and discover the trend in atomic radius in a group. Then, have some groups present their findings to the class. After they complete the presentations, harmonize the concepts with the facts as follows:

In going down a group, the atomic radius of the elements increases, due to a progressive increase in the number of shells. Thus, for Group IA metals, atomic size increases down the group from lithium to cesium, as shown in Table 2.1.

Table 2.1 Atomic radii and number of shells of Group IA metals.

Element	Number of Shells	Atomic Radius (Å)
Li	2	1.34
Na	3	1.54
K	4	1.96
Rb	5	2.16
Cs	6	2.35

Then, proceed by presenting ionization energy. First, ask them to define ionization energy. After getting their responses, give them the correct definition of ionization energy and also explain the difference between first and second ionization energy. Tell them what ionization energy measures.

Next, have students discuss Activity 2.7 in groups for a few minutes and identify the trend in ionization energy of elements in a group. Have one or two groups present their findings to the class. After they complete their presentations, harmonize concepts, using the following information.

Generally, with increasing atomic number, the first ionization energy of the elements decreases down a group. The following tables illustrate the first ionization energy of alkaline earth metals, alkali metals and halogens.

Table 2.2 Ionization energy of alkaline earth metals

Element	Ionization energy (kJ/mol)
Be	899
Mg	738
Ca	590

Table 2.3 Ionization energy of alkali metals

Element	Ionization energy (kJ/mol)
Li	520
Na	496
K	419

Table 2.4 Ionization energy of halogens

Element	Ionization energy (kJ/mol)
F	1681
Cl	1251
Br	1140

After that, tell the students about the factors affecting ionization energy.

Activity 2.8 is designed to help students to relate an element and its properties. So, have them discuss this activity for a few minutes in groups. Encourage one or two groups to present their conclusions to the class. Following the presentations, harmonize concepts using the following information.

Element X - From the given information, element X is found in period 3 and Group VI. Thus, the element is sulphur (S) with atomic number 16.

Element M - is characterized by low ionization energy, forming an oxide with the formula M_2O. It must be a metal with one valence electron. Since element M is an alkali metal found in the fourth period, the element M is potassium (K).

Element Y- A closer look at the configuration of transition elements helps students to see that element Y has atomic number 30. This is because it has 10 electrons in the 3*d* orbital after filling the 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* sublevels, which accommodate 20 electrons. This element is zinc, and its configuration is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}$. It is a coinage metal in Ethiopia.

After introducing ionization energy, continue with electron affinity. First, ask the students to define **electron affinity**. After their responses, define it yourself and tell them what it measures. Next, have them discuss **Activity 2.9** in groups for a few minutes and then ask some groups present their opinions to the class. Following the presentations, harmonize concepts, using the following information:

Electron affinity depends on the size of the atom and effective nuclear charge.

1. Down a given group, even though nuclear charge increases, the effective attraction of the nucleus for an added electron decreases. This is due to the increasing shielding effect of the inner electrons. Therefore, electron affinity decreases down a group. The following table shows the electron affinity of the halogens.

Table 2.5 Electron affinities of the halogens

Elements	Electron affinity (kJ/mol)
F	−328
Cl	−349
Br	−325
I	−295

The electron affinity of chlorine is higher then that of fluorin, due to smaller size of fluorine. The smaller the size of the atom, the greater the force of repulsion between the electron being added to the atom and the electrons already present on the atom, so that the smaller the electron affinity of the element would be,

2. Noble gases have extremely low (almost *zero*) electron affinity, due to their stable electron configuration and low tendency to accept an additional electron.
3. Halogens have the highest electron affinities (*highest negative values*). Halogens need only one electron to be stable. Therefore, they have a high tendency to gain an electron and form stable negative ions.

Next, continue by introducing electronegativity. First, ask them to define electronegativity. After their attempt, tell them the correct definition of electronegativity and inform them of its relationship with ionization energy and electron affinity. Have them discuss **Activity 2.10** in groups for a few minutes and discover the trend in electronegativity. Invite one or two groups to present their findings. After the presentations, harmonize concepts, using the following information:

Electronegativity decreases down a given group, due to increase in atomic size. Generally, high electronegativity is characteristic of nonmetals, and low electronegativity is characteristic of metals.

The decreasing order of electronegativity for each set of elements is:

a. Be > Mg > Ca > Ba

b. C > Si > Pb > Ge

c. F > Cl > Br > I

The following tables show the electrnegativity values for group IIA, group IVA and group VIIA elements.

Table 2.6 Electronegativity of alkaline earth metals

Element	Electronegativity
Be	1.5
Mg	1.2
Ca	1.0
Ba	0.9

Table 2.7 Electronegativity of halogens

Element	Electronegativity
F	4.0
Cl	3.0
Br	2.8
I	2.5

Table 2.8 Electronegativity of Group IV elements

Element	Electronegativity
C	2.55
Si	1.90
Ge	1.96
Pb	2.01

After presenting electronegativity, introduce the trend in metallic character of the elements in a given group. Explain the relationship between metallic character and atomic size. Tell them the positions of elements in the periodic table, in terms of their strength in metallic character.

Then, have students discuss **Activity 2.11** in their groups for a few minutes to identify the trend in metallic character of elements down group IV. Then ask some groups to present their findings to the class. After the presentations, harmonize concepts suggested by the students with the truth, using the following information:

1. Group-IVA elements have metallic, nonmetallic, and metalloid characteristics, as shown in the following table.

Table 2.9 Metallic character of Group-IVA elements

Elements	Character of the element
C	non metal
Si	metalloid
Ge	metalloid
Sn	metal
Pb	metal

2. Atomic size :- Pb > Si

   Ionization energy:- Si > Pb

   Electron affinity :- Si > Pb

   Electronegativity :- Si > Pb

Period properties within a period

To teach this subtopic, it is advisable to use group discussion as your methodology. However, you can create a gapped-lecture where you need one.

The topic starts with an activity. Activity 2.12 is designed to help students identify the positions of the most active metals and the most active nonmetals in the periodic table. Have the students discuss this activity in groups for a few minutes. Then ask some groups to present their findings to the class. After they complete their presentations, tell them that:

1. Metals are found on the left side of the periodic table, and nonmetals are on the right side. Metalloids are found between metals and nonmetals.
2. The most active metals are found in the left bottom corner, and the most active nonmetals are at the top right corner of the periodic table.

After harmonizing concepts related to Activity 2.12, continue by presenting periodic properties of elements across a period. First, treat atomic size by having students discuss **Activity 2.13** in groups for a few minutes. Encourage some groups to present their findings to the class. After the presentations, harmonize concepts as follows:

From left to right in a given period, nuclear charge or atomic number increases progressively, by one, for every succeeding element, as an increasing number of valence electrons is added to the same shell. This results in an increase in effective nuclear charge. Due to this, the valence electrons are pulled closer to the nucleus. As a result, atomic size of the elements decreases across a period.

Table 2.10 Atomic radii of Periods 2 Elements

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Atomic radii (Å)	1.34	0.90	0.82	0.77	0.75	0.73	0.71	0.69

Then, continue by presenting ionization energy. First, have the students recall what factors affecting it's ionization energy. After that, have them discuss **Activity 2.14** in groups for a few minutes and then ask some groups present their findings. After the presentations, harmonize concepts, using the following information:

1. Across a period, ionization energy increases. Two factors account for this effect. First, across a given period, nuclear charge increases steadily from left to right. Also, the nucleus increasingly attracts each additional valence electron which, in effect, increases the effective nuclear charge and decreases the atomic size. These results cause increase in ionization energy from Na to Ar.
2. However, for various reasons, some irregularities are observed across a period.

   i. Al - $1s^2 2s^2 2p^6 \underline{\mathbf{3s^2 3p^1}}$ and Mg - $1s^2 2s^2 2p^6 \underline{\mathbf{3s^2}}$

   The first ionization energy of Mg > Al. This is due to **completely-filled sublevel stability**.

   ii. S - $1s^2 2s^2 2p^6 \underline{\mathbf{3s^2 3p^4}}$ and P - $1s^2 2s^2 2p^6 \underline{\mathbf{3s^2 3p^3}}$

   The ionization energy of S < P. This is because of **half-filled sublevel stability** in the case of phosphorous, P.

Tell your students that it is difficult to remove an electron from an atom possessing half-filled or completely filled outermost subshell, and hence such atoms have ionization energies.

Next, introduce the trend in electron affinity across a given period and continue with the trend in electronegativity. Before you deal with present the details, have students discuss Activity 2.15 in groups for a few minutes to discover the trend in electronegativity of elements across a given period. Then ask some groups present their findings. After the presentations, continue by harmonizing concepts as follows:

1. The following table shows the electronegativity values of Period 3 elements.

Table 2.11 Electronegativity values of Period-3 elements

Period-3 element	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

2. Generally, across a period, electronegativity increases because of the gradual increase in effective nuclear charge and also the decrease in atomic size.

Finally, inform them about the trend in metallic character from left to right across a given period. At the end of the section, have them do the suggested project work.

Project work: Students should be encouraged to use locally available materials make this model of a section of the periodic table. Using their model, they are also expected in to fill this section of the periodic table with the appropriate elements position. Students should think of reasons for changes in size and the patterns from Li to F, Na to Cl, and from Li to Cs, using the data in the table.

$_{3}Li$	$_{4}Be$		$_{5}B$	$_{6}C$	$_{7}N$	$_{8}O$	$_{9}F$	$_{10}Ne$
$_{11}Na$	$_{12}Mg$		$_{13}Al$	$_{14}Si$	$_{15}P$	$_{16}S$	$_{17}Cl$	$_{18}Ar$
$_{19}K$								
$_{37}Rb$								
$_{55}Cs$								

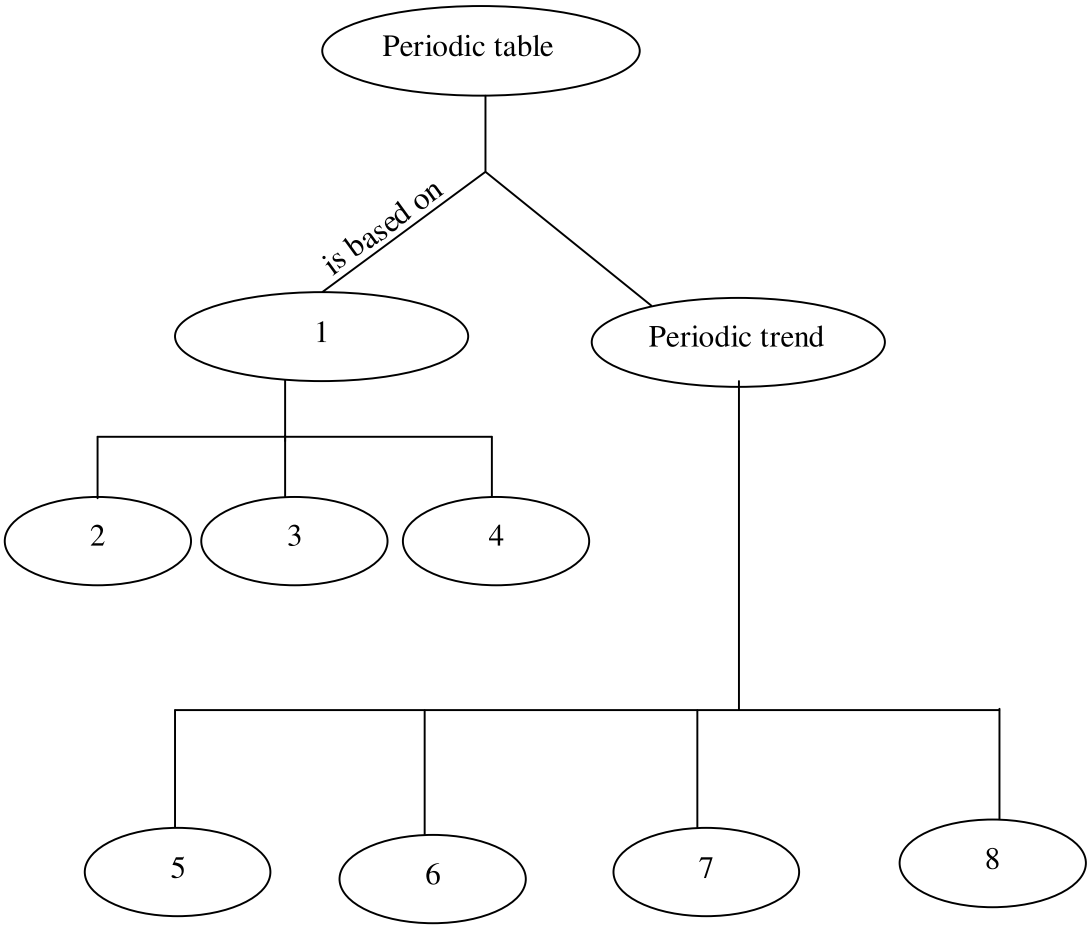
i. For the decrease in size from Li to F and Na to Cl, students should understand that it is due to an increase in the nuclear charge attraction for the extra added electrons filling the same valence shell.

ii. The increase in size from Li to Cs is accounted for by the increase in the number of energy levels.

Concept Mapping

After you introduce the classification of the elements (the periodic table), list the following terms on the blackboard and let some volunteer students try to complete the concept map using the terms:

- electronegativity
- electron configuration
- atomic radius
- group number
- ionization energy
- period number
- block
- electron affinity



Appreciate their attempts in constructing the concept map and harmonize by drawing the concept map on the blackboard. Let the students compare the concept map with their attempts.

1. Electron configuration
2. Group number
3. Period number
4. Block
5. Electronegativity
6. Ionization Energy
7. Electron Affinity
8. Atomic radius

Assessment

You are advised to assess each student's work throughout Section 2.3. This can be done by recording the performance of every student. You can record the performance of each student by taking into consideration how the student:

- takes part in discussing Activities 2.4 – 2.15.
- involves in presenting concepts after discussion.
- answers questions raised during: discussions, harmonizing concepts, stabilization and gapped lectures.
- attempts to construct the concept map suggested in this section.

Give them Exercises 2.4 and 2.5 as class work or homework. Check their work and record their performances.

From the record, make sure that the suggested competencies for the section are achieved by most of your students. Praise students working above the minimum requirement level and recognize their achievements. In case of students working below the minimum requirement level, give them assistance in class and arrange additional lesson time. You can also give them exercise from the additional questions recommended to them and prepare others related to the contents in this section. This will help them to catch up with the rest of the class.

Additional Questions

1. Although carbon and lead are in the same group, carbon is a nonmetal, and lead is a metal. Explain how two elements with the same number of valence electrons can have such different properties.

* 2. How many inner electrons and valence electrons are present in an atom of the following elements?
   i. Ca
   ii. Br
   iii. O
   iv. Cs
   v. Kr

* 3. Which element is described by each of the following descriptions? Place the correct letter of the question (a, b, c, etc) in the box given below.
   a. Smallest atomic radius in Group VIA
   b. Largest atomic radius in Period 5
   c. Highest first ionization energy in Group IA
   d. Most metallic in Group IIIA

e. The transition metal with the lowest atomic number

f. Extremely low election affinity in Period 3

g. The element with the electron configuration [Kr] $5s^2 4d^6$

h. Period-4 element with a completely filled outer energy level.

i. The most electronegative element

j. Period-5 transition element with the lowest ionization energy

Answers to Additional Questions

1. Generally atomic size progressively increases down a group. Thus the atomic size of lead is much greater than that of carbon. Since with increasing atomic size nonmetallic behavior of the elements decreases, lead is a metal whereas carbon is a nonmetal.

2. First, write the electron configurations of each element and then you can determine the number of inner shell electrons and valence electrons.

	Element	Inner shell electrons	Valence electrons
i	Ca	18	2
ii	Br	28	7
iii	O	2	6
iv	Cs	54	1
v	Kr	28	8

3.

c																	
															a	i	
																	f
		e															h
b		j					g										
												d					

Answers to Exercises

Exercise 2.4

1. D 2. B 3. C

4. The valence electrons are shielded or screened from the full attraction of the nucleus by the inner electrons and this phenomenon is known as **screening effect.**

   Effective nuclear charge, Z_{eff}, is the difference between the actual nuclear charge (Z) and the screening by inner electrons (S).

   $$Z_{eff} = Z - S$$

5. Elements having low ionization energy are more metallic.

Exercise 2.5

1. D 2. A 3. C

Critical Thinking:

4 Students should be helped to see the causes of the electronic configurations of the three elements.

$C – 1s^22s^22p^2$

$N – 1s^22s^22p^3$

$O – 1s^22s^22p^4$

Contrary to the expected trend, N has higher ionization energy than O and C do, due to its half-filled sublevel ($2p^3$) stability. Thus, the trend is N > O > C.

5 Consider the sizes and configurations of the two elements. Fluorine has two energy levels (2, 7) and chlorine has three energy levels (2, 8, 7). The extremely small size of fluorine causes strong electron-electron repulsion when an extra electron is added. Therefore, the electron affinity of fluorine is lower than that of chlorine, even though fluorine might be expected to have a higher electron affinity. The electronegativity value of fluorine is, however, higher due to the stronger tendency to attract electrons.

2.4 Advantages of Periodic Classification

Periods Allotted: 1 period

Competencies

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

- describe the advantages of periodic classification in chemistry.

Forward Planning

Make the necessary preparation to teach the concepts on the advantages of periodic classification by reading the contents thoroughly. Read the teacher's guide to get information about the methodologies you need to implement and to gain more ideas about the suggested activity (Activity 2.16). Prepare a detailed plan that shows the time allotted for every activity you perform during the period. You also need to have a big periodic table chart for use during the teaching-learning process.

Teaching Aids

- Periodic table chart.

Subject Matter Presentation

You better use group discussion, visual-based active learning and brainstorming methodologies.

For this section, you are advised to use group discussion, brainstorming and visual-based learning methodologies. After introducing the topic of the lesson, let the students discuss Activity 2.16 in groups for a few minutes. With this activity, students should conclude their study of the periodic table by forming general ideas about the uses of the periodic table. Here apply brain storming methodology. Encourage some groups to present their opinions. Record all suggested ideas of the students on the blackboard. Then, continue harmonizing concepts by listing the advantages of periodic classification of the elements and the information that can be obtained from the periodic table. Also tell them about diagonal relationship, the elements showing this relationship and the pe riod numbers where the elements are found.

Assessment

Assess each student's work throughout the section. You can do this by using student's performance list. Record how every student:

- involves in group discussion.
- participates in presenting opinions of the groups after discussion.
- suggests ideas while you implement brain storming method.
- gives comments on wrong views suggested by students.
- answers questions raised during harmonizing concepts.

You can also prepare questions related to advantages of periodic classification and give them as class work or homework. Check their work and record their performances. From the cumulative records, make sure whether or not the suggested competencies for this section are achieved. Appreciate students working above the minimum requirement level and help those working below the minimum requirement level. Let the students also attempt the review exercise on unit 2.

Answers to the Review Exercise of Unit 2

Part I	**Part II**
1. Wrong	8. B
2. Wrong	9. B
3. Wrong	10. D
4. Wrong	11. B
5. Correct	12. B
6. Wrong	13. A
7. Correct	

Part III

	Group	Period	Block type
14	IIA	4	*s*
15	VIA	3	*p*
16	VIIIB	4	*d*
17	IIB	4	*d*
18	VIIIA	5	*p*

Part IV

19. This is the close relationship between elements in neighboring groups of the periodic table. Li, Be and B have more in common with the Period-3 elements in the next group than with the Period-3 element in their own group (Li with Mg, Be with Al, and B with Si). These sets of elements have similar chemical properties.

20. a. Br has the highest first ionization energy because it is nonmetallic.

    b. Na has the smallest atomic size; two energy levels with only 11 positive charges in the nucleus.

    c. Na is most metallic.

21. N - $1s^2 2s^2 2p^3$ (2, 5)   O - $1s^2 2s^2 2p^4$ (2, 6)

    Because oxygen has a higher effective nuclear charge, it is more electronegative.

22. Na - $1s^2 2s^2 2p^6 3s^1$ (2, 8, 1)   Cl - $1s^2 2s^2 2p^6 3s^2 3p^5$ (2, 8, 7)

Though they have the same number of energy levels, the effective nuclear charge is higher in Cl, resulting in a small atomic size.

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Website

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http://en.wikipedi.org/wiki/periodic-table.

Unit 3 Chemical Bonding and Intermolecular Forces

Unit Overview   Total Period Allotted 17

In this unit, students will learn how atoms bond to each other to form molecular compounds, and will be able to understand the attractive forces between molecules.

Unit 3 has five main parts:

- **Section 3.1**, presents chemical bonding, electron-dot diagrams (also called Lewis structures), and valence electrons.
- The basic goal of **Section 3.2** is to explore the formation of ionic bonds. It deals with the ionic bond and its formation by the transfer of electrons. This section illustrates the formation of ionic bonds between metals and nonmetals, using Lewis structures.
- **Section 3.3** deals with covalent bonding: The single bond, double bond and triple bond are described. The formation of each is illustrated by Lewis structures.
- **Section 3.4** illustrates the formation of metallic bonds. It illustrates how free electrons surround the positive ions of metal and explains the resulting metallic bonding.
- The last section, **Section 3.5,** explains intermolecular forces, including dipole-dipole forces and hydrogen bonds. This section particularly emphasizes to the physical properties of substances, such as boiling point.

In general, this unit emphasizes the formation of compounds. It helps the students to develop the skill of drawing electron structures.

You can use group discussion, gapped lecture, question and answer, visual-based active learning, experiment as your methods of teaching for this unit.

Unit Outcomes

After completing this unit, students will be able to:

- discuss the formation of ionic, covalent and metallic bonds;
- know the general properties of substances containing ionic, covalent and metallic bonds;

- develop the skills of drawing the electron dot or Lewis structures for simple ionic and covalent compounds;
- understand the origin of polarity within molecules;
- understand the formation and nature of intermolecular forces;
- appreciate the importance of intermolecular forces in plant and animal life;
- demonstrate scientific inquiry skills: observing, predicting, making model, communicating, asking questions, measuring, applying concepts, comparing and contrasting, relating cause and effects.

Main Contents

3.1 Chemical Bonding

3.2 Ionic Bonding

3.3 Covalent Bonding

3.4 Metallic Bonding

3.5 Intermolecular Forces

Answers to Review Exercises

3.1 Chemical Bonding

Period Allotted 1

Competencies

After completing this section, students will be able to:

- define chemical bonding;
- explain why atoms form bonds.

Forward Planning

Prepare yourself on the general concept of chemical bonding by reading the student's text, reference books and other resources. Also read the teacher's guide to get more information about the start-up activity and Activity 3.1 and 3.2 as well as the methodologies you implement.

Subject Matter Presentation

You can use group discussion and gapped lecture as your methods of teaching for this section.

You are advised to begin the unit with the start-up activity given in the student's textbook. Have students discuss in groups and share their ideas with the class.

The startup activity introduces students to the nature and strength of the forces between the atoms in forming a bond. After introducing the unit, the teacher is advised to inform students the kind and strength of bonds between atoms. Emphasis should be given only to covalent bonds at this stage. The three set of sticks show that there are three types of covalent bonds – single, double and triple bonds. When students see the sets of sticks being broken, they should conclude that a triple bond is stronger than a double bond, and a double bond is stronger than a single bond.

Start the lesson on this section using Activity 3.1. The activity enables students discover the cause for the chemical combination of elements to form compounds. So, have students discuss Activity 3.1 in groups for a few minutes. Then, have some students from different groups present their findings to the class. After their presentations, harmonize the discussion by explaining the following concepts.

Most elements are not found free in nature because they are unstable in the free state and combine with one another to form molecules or compounds. In the combined state (form) they become more stable than in their free state.

A chemical bond is the force of attraction between two or more atoms or ions in molecules or compounds. Emphasize how atoms form bonds in order to achieve more stable electronic arrangements similar to those of the noble gases.

Atoms can attain a stable electron arrangement by losing, gaining or sharing electrons. When atoms lose, gain or share electrons, they also attain lower potential energy states, which results in greater stability.

After introducing why atoms form chemical bonds, ask students whether atoms of all elements readily form chemical bonds or not. Also, ask them which atoms have unstable and which atoms have stable electron configurations, on the basis of valence shell electron configurations.

Next, proceed to Activity 3.2 This activity is designed to help students realize why noble gases are stable and also understand why and how most elements tend to achieve electrons configuration similar to that of the noble gases. Therefore, have students perform Activity 3.2 in groups for a few minutes and then ask some groups to present their findings to the class. Write the table filled with electron configurations on the blackboard and help students to conclude what is common about all these elements (except He). Ask them whether they are stable or not.

Finally, harmonize the discussion by presenting the following information. These Group-VIIIA elements (known as noble gases) have stable ns^2np^6 outer shell electron configurations or 8 electrons in their valence shell, except for helium. For He, s^2 is completely filled and stable. Atoms lose, gain, or share electrons to attain stable electron configurations of the noble gases.

Table 3.1 Valence electrons of the noble gases

Elements	Atomic Number	Valence-Shell Electronic Configuration	Number of Valence Electrons
Helium, He	2	2	2
Neon, Ne	10	2,8	8
Argon, Ar	18	2,8,8	8
Krypton, Kr	36	2,18,8,8	8
Xenon, Xe	54	2,8,18,18,8	8
Radon, Rn	86	2,8,18,32,18,8	8

Helium has 1 energy level occupied by 2 electrons. This gives helium a stable noble-gas configuration, and therefore it is placed in GroupVIIIA. After harmonizing concepts, state the octet rule and tell the students that valance electrons are responsible for the formation of chemical bonds.

Assessment

You can assess each student's work throughout this section by recording his/her performance. You can record the performance by considering how the student:

- involves in group discussion (Activity 3.1 – 3.2 and start-up activity).
- takes part in presentations after discussion.
- answers questions raised during harmonizing concepts and gapped lectures.

You can also use the additional questions given in the guide and other questions of your own as class work or homework. Check their work and record their performances. Based on the record you have, see whether or not the competencies suggested for this section are achieved.

Praise students who are working above the minimum requirement level and recognize their achievements. Give the necessary support to students working below the minimum requirement level either by letting them do additional exercises or by arranging extra lesson time.

Additional Questions

* 1. Which family of elements is extremely unreactive under ordinary conditions? Explain your answer.

* 2. Which are more stable under ordinary conditions, hydrogen atoms or hydrogen molecules? Explain.

3. Which atom has higher potential energy, argon or chlorine? Why?

Answers to Additional Questions

1. Group VIIIA elements or noble gases have 8 valence electrons (except He) and thus they are chemically stable. Because of their stable electron configuration they do not need to react with other elements.
2. Hydrogen molecule (H_2) is more stable than hydrogen atom. This is because hydrogen atom needs more electrons to be stable or to achieve the helium electron configuration.
3. Chlorine atoms are at relatively high potential energy because they exist in a less stable state.

Answer to Exercise 3.1

1. Many atoms are less stable when they exist free than when they are combined. Hence, when atoms combine to form compounds, they will attain the lowest energy states and became stable.
2. A chemical bond is formed by **losing** and **gaining** (transferring) or **sharing of electrons**.
3. Electrons in the outermost shell of an atom take part in a chemical bond.
4. Halogens have 7 valence electrons, whereas noble gases have 8 valence electrons. Halogens tend to gain one electron in order to be stable, whereas noble gases have already filled 8 valence electrons. Hence, halogens are reactive whereas noble gases are unreactive.

3.2 Ionic Bonding

Period Allotted 3

Competencies

After completing this section, students will be able to:

- explain the term ions;
- illustrate the formation of ions by giving examples;
- define ionic bonding;
- describe the formation of an ionic bond;
- give examples of simple ionic compounds;
- draw Lewis structures or electron dot formulas of simple ionic compounds;
- explain the general properties of ionic compounds;
- investigate the properties of a given sample of ionic compounds.

Forward Planning

Make a big chart that illustrates the formation of ions and Lewis structures of simple compounds. Prepare yourself by reading the main concepts about ionic bonding, such as ionic bond formation and general properties of ionic compounds from the students' text book. Check whether or not materials and chemicals are available in the laboratory for the experiment given in this subunit. Try the experiment before the lesson.

Teaching Aids

- Chart showing formation of ions and Lewis structures of simple compounds.
- Refer to the students' text for the chemical and apparatus required to perform experiment 3.1.

Subject Matter Presentation

Ionic bonding

You are advised to use group discussion, question and answer and visual-based active learning and experiment as methods of teaching. This section begins with an activity.

Activity 3.3 tells students why metals lose and nonmetals gain electrons easily. Have the students discuss activity 3.3 in groups and then present their opinions to the class. Harmonize the discussion, using the following facts.

1. Metals lose electrons easily. This is because metals have 1, 2 or 3 valence electrons, these valence electrons are loosely bound, and their ionization energy is low.
2. NaCl in solution and in molten state dissociates into Na^+ and Cl^- ions, but in solid state it forms a crystal. NaCl conducts electric current when dissociated into ions and does not conduct in crystalline state.

After harmonizing concepts, tell students how atoms form anions and cations and the ionic charge they acquire. Then, continue with Activity 3.4. The activity enables students to identify the atoms that lose or gain electrons from their electron configuration. So, let them discuss Activity 3.4 in groups for a few minutes and then have some groups present their findings to the class.

After the presentations, harmonize concepts as follows. Metals in Group IIA, alkaline earth metals, tend to lose two electrons when they combine with other elements, forming ions of +2 charge. Ca and Ba lose two electrons to form cations of +2 charge.

$$Ca \longrightarrow Ca^{2+} + 2e^-$$

$Ba \longrightarrow Ba^{2+} + 2e^-$

Group VIA elements, Chalcogens (*ore forming elements*), accept two electrons and form anions with –2 charge. O and S each accept two electrons to form ions with –2 charge.

$O + 2e^- \longrightarrow O^{2-}$

$S + 2e^- \longrightarrow S^{2-}$

Students should be able to relate the position of the elements in the periodic table to the normal ion formed by them. For example

Group 1 -	M^+,	e.g. Li^+, Na^+, K^+
Group 2 -	M^{2+},	e.g. Be^{2+}, Mg^{2+}, Ca^{2+}
Group 3 -	M^{3+},	e.g. Al^{3+}
Group 5 -	M^{3-},	e.g. N^{3-}, P^{3-}
Group 6 -	M^{2-},	e.g. O^{2-}, S^{2-}
Group 7 -	M^-,	e.g. F^-, Cl^-, Br^-

After they gain this background knowledge, enable the students to write the symbols of ions formed by atoms of different elements.

Then, proceed to introduce ionic bond formation. After getting responses, tell them the type of atoms that combine to form ionic bonds and define what an ionic bond is. Show them how the transfer of electrons occurs during the formation of ionic bonds, using Bohr diagrams or electron-dot notation or Lewis symbols of the combining atoms. Have the students practice drawing Bohr's diagrammatic representation and writing Lewis formulas for some ionic compounds.

Use Activity 3.5 for this purpose. Have them do this activity in groups for a few minutes. After that, encourage one or two students to perform the task on the following information. Introduce them to naming ionic compounds. Make sure that students can show the formation of ionic bonds using either diagrams or Lewis symbols and they continue with the properties of ionic compounds.

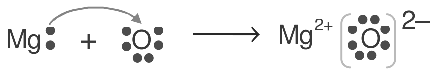
Show the students Bohr's diagrammatic representation by drawing the atoms and the compounds on a big piece of chart paper after their presentation.

In general, an ionic bond is formed by the transfer of electron(s) from a metal to a nonmetal.

Metal + Nonmetal ⟶ Ionic compound

a. Potassium chloride ⇒

b. Magnesium oxide ⇒



c. Calcium chloride ⇒

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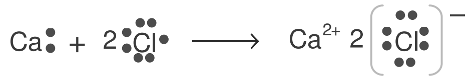
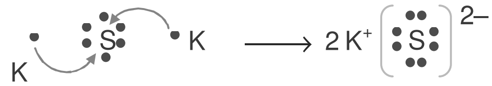


d. Potassium sulphide ⇒



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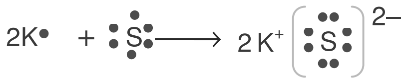


General Properties of Ionic Compounds

To teach this topic, use group discussion and brain storming as your methodologies. To apply the brain storming methodology, ask the students to tell you the properties of ionic compounds either in groups or individually. Record the properties suggested by the students on the blackboard. After that, harmonize concepts by giving corrections and letting them discover those properties that do not describe ionic compounds.

Next, give the students Activity 3.6 to to perform in groups. Guide them in collecting samples of ionic compounds from the school laboratory. Have the students bring some of these samples to the classroom and then describe the physical properties of the samples to the class. The properties of ionic compounds listed in the student's text book will help students to identify and characterize ionic compounds.

Harmonize the discussion by relating the actual properties with those of the substances presented to the class.

Following this, let the students do Experiment 3.1 in groups, as given in the student's textbook. Have the students write what they observe during the experiment and report their findings to the class.

Experiment 3.1

This experiment helps students understand the physical properties of ionic compounds, such as melting point, solubility and conductivity.

I. Melting point and solubility

First of all, remind the students that both NaCl and $CuCl_2$ are ionic compounds. Ionic compounds have high melting points and are soluble in polar solvents like water. Based on these facts, note the following observations from the experiments:

In Experiment 1, neither NaCl nor $CuCl_2$ melt when heated with the flame of a Bunsen burner. This is because the heat produced by a Bunsen burner is too low to melt either NaCl or $CuCl_2$ crystals. The melting point of NaCl is 801°C, and the melting point of $CuCl_2$ is 498 °C.

a) In Experiment 2, both NaCl and $CaCl_2$ dissolve in water, but not in organic solvents like ethanol, hexane and benzene.

Substance	Water	Ethanol	Hexane	Benzene
NaCl(s)	Soluble	Insoluble	Insoluble	Insoluble
$CuCl_2$	Soluble	»	»	»

II. Conductivity

From this experiment, students learn the reason for the electrical conductivity of compounds in aqueous solution. Set up the apparatus as shown in the diagram.

Help students to see the conductivity for solid NaCl, then for a solution of NaCl after dissolving it. Replace the NaCl solution with a solution of copper (II) chloride. Have the students write down their observations. Continue to check whether or not benzene and charcoal conduct electricity in water.

Have your students analyze their observations as to why sodium chloride does not conduct electric current in solid state. The bulb lights when the NaCl or $CuCl_2$ is dissolved. Ask them what process is taking place in the dissociation of NaCl and $CuCl_2$. Help them to conclude that ionic compounds conduct electric current in aqueous solution due to their dissociation to positive and negative ions. Finally, ask them why benzene and charcoal do not conduct electric current, even in solution. Note that benzene and charcoal contain only covalent bonds, and do not produce anions and cations.

Assessment

Assess how every student is doing throughout section 3.2. You need to follow-up how every student:

- takes part in discussing Activity 3.3 – 3.6
- involves in presentations after discussion
- participates in performing Experiment 3.1

- involves in presenting observations about the experiment
- answers questions raised during discussions, harmonizing concepts, and stabilization and record their performances in the students' performance list.

You can also give them Exercise 3.2 and 3.3 and other questions of your own relevant to this section, as class work or homework. Check how well they have done and record their achievements. From the cumulative record you have, see whether or not the students have achieved the competencies suggested for the section. Appreciate students working above the minimum requirement level and encourage them to continue working hard. To assist students working below the minimum requirements level, catch up with the rest of the class, give them additional exercise related to the formation of ionic bonds and properties of ionic compounds or arrange extra lesson time.

Additional Questions

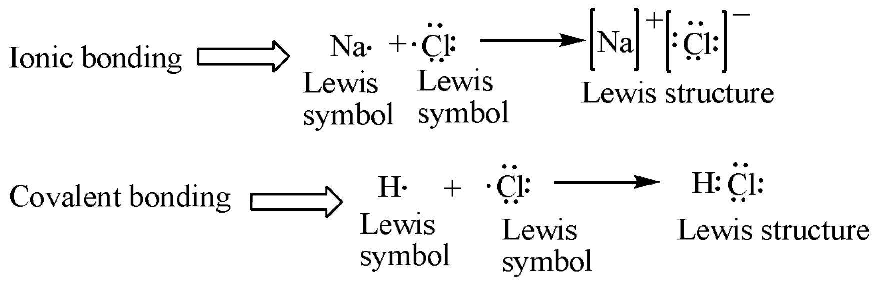
1. What is the difference between a Lewis symbol and a Lewis structure?

* 2. Predict the change that must occur in the electron configuration if each of the following atoms is to achieve a noble gas configuration.
   - i. Phosphorous
   - ii. Potassium
   - iii. Aluminium
   - iv. Bromine

* 3. Draw the Lewis symbol for each of the following elements.
   - i. Silicon
   - ii. Alumnium
   - iii. Krypton
   - iv. Calcium
   - v. Bromine

Answers to Additional Questions

1. **A Lewis symbol** consists of a chemical symbol together with dots that are placed around the symbol. The chemical symbol represents the atom and the dots represent the valence electrons.
   A Lewis structure is a combination of Lewis symbols that represents either the transfer or the sharing of electrons in a chemical bond.
   The difference between a Lewis symbol and a Lewis structure can be illustrated using the following two examples:



2. Generally, metals lose electrons and nonmetals gain electrons to achieve the electron configuration of the nearest noble-gas element. Therefore, the electron configuration of P, K, Zn and Br after achieving the noble gas configuration can be written as follows:

(i) Phosphorous (atomic no = 15) gains 3 electrons to achieve the electron configuration of argon.

$$P^{-3} = 18\,\text{electons}; 1s^2 2s^2 2p^6 3s^2 3p^6 = 2,8,8$$

(ii) Potassium (atomic no = 19) loses 1 electron so that it becomes stable.

$$K^{+} = 18\,\text{electrons}; 1s^2 2s^2 2p^6 3s^2 3p^6 = 2,8,8$$

(iii) Aluminum (atomic no = 13) loses 3 electrons and achieve the electron configuration of neon.

$$Al^{+3} = 10\,\text{electrons}; 1s^2 2s^2 2p^6 = 2,8$$

(iv) Bromine (atomic no = 35)

$$Br^{-} = 36\,\text{electrons}; 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^0 4p^6 = 2,8,18,8$$

3. i. Silcon ⟹ ·Si·
ii. Aluminium ⟹ ·Al·
iii. Krypton ⟹ :Kr:
iv. Calcium ⟹ •Ca•
v. Bromine ⟹ :Br·

Answers to Exercises

Exercise 3.2

1 The designation for nitrogen, N, is $^{14}_{7}N$.

Its electron configuration is 2, 5.

∴ The outermost shell of nitrogen contains five (5) electrons.

Its Lewis structure is $\cdot\ddot{N}\cdot$ (with one dot below)

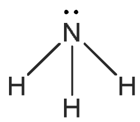
The Lewis structure for the nitrogen molecule, N_2, is

$:N⋮⋮N:$ or $:N \equiv N:$

The Lewis structure for ammonia, NH_3, is

H:N̈:H
H

or



2 The electron configuration for each species is given as:

$_{11}Na^{+} = 2, 8$ ………………. *by losing one electron.*

$_{12}Mg^{2+} = 2, 8$ ……………..... *by losing two electrons.*

$_{10}O^{2-} = 2, 8$ ………………. *by gaining two electrons.*

$_{10}Ne = 2, 8$

$\therefore$ $_{11}Na^{+}$, $_{12}Mg^{2+}$, $_{8}O^{2-}$, and $_{10}Ne$ have the same number of electrons.

Exercise 3.3

1. KCl is an ionic compound and can dissolve only in polar solvents like water. This can be generalized by the solubility principle '*like dissolves like*'. KCl is insoluble in benzene because benzene is a nonpolar solvent.
2. NaCl(aq) and NaCl(ℓ) conduct electric current because these compounds dissociate into ions (Na^{+} and Cl^{-}) which can carry electric charge. But NaCl is a poor conductor in the solid state because it does not dissociate into ions.
3. a. Calcium sulphide
   b. Sodium iodide
   c. Silver bromide
4. Ionic compounds
   - are composed of oppositely charged ions but do not contain molecules.
   - are hard and rigid crystalline solids at room temperature.
   - have relatively high melting and boiling points.
   - can conduct electric current when molten or in aqueous solution.
   - are usually soluble in polar solvents.

3.3 Covalent Bonding

Period Allotted 8

Competencies

After completing this section, students will be able to:

- define covalent bonds;
- describe formation of covalent bonds;
- draw Lewis structures or electron-dot formulas of simple covalent bonds;
- give examples of different types of covalent bonds;
- make models of covalent bonds to show single, double and triple bonds, using sticks and balls or other locally available materials;
- explain polarity in covalent bonds;
- distinguish between polar and nonpolar covalent molecules;
- define coordinate (dative) covalent bond;
- illustrate the formation of coordinate covalent bonds using appropriate examples;
- explain the general properties of covalent compounds;
- investigate the properties of given samples of covalent compounds.

Forward Planning

Before the lesson, read about the main concepts of covalent bonding such as the formation of covalent bonding, Lewis formulas of covalent molecules, polarity, coordinate covalent bonds and the general properties of covalent compounds. Use the students' textbook, reference books and other resources. Check whether or not the materials and chemicals are available in the laboratory for Experiment 3.2 given in this subunit. Carry out the experiment yourself before the lesson. Make a big chart and models that illustrate the formation of covalent bonds, using Lewis formulas.

Teaching Aids

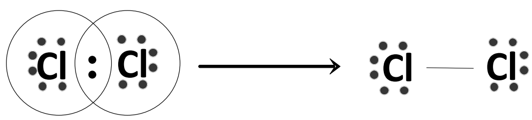
- charts and models
- Refer to the students' text for the chemical and apparatus required to perform experiment 3.2.

Subject Matter Presentation

Use group discussion and question and answer and experiment methodologies to teach this topic. Begin the lesson with Activity 3.7 as given in the student's text book. The experiment helps students to differentiate the formations of ionic and covalent bonds and also their properties. Let them discuss this activity for a few minutes. Then have them present the ideas of their discussion to the class. Also, allow them to explain according to their understanding.

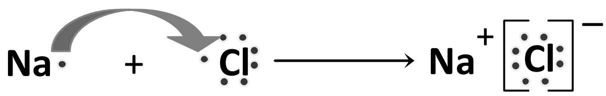
Next, harmonize the students' discussion by giving the answer (s) as presented below, if needed.

1. Two chlorine atoms combine to form a chlorine molecule by sharing electrons between them.



Such a bond formed by the sharing of electrons is called a **covalent bond**.

In the second case, sodium combines with chlorine to form sodium chloride by the transfer of an electron from the outer most shell of sodium to chlorine.



This type of a bond is called an ionic bond. Therefore, chlorine in Cl_2 forms a covalent bond, whereas in NaCl, the bond is ionic.

2. Carbon tetrachloride is a nonpolar molecule, and it does not dissociate into ions. Therefore,
   i) CCl_4 is a poor conductor of electricity.
   ii) CCl_4 is insoluble in polar solvents like water, but is soluble in nonpolar solvents like benzene and ether.

Next, have the students understand the formation of covalent bonds in HCl, F_2, CH_4 and H_2O. Tell them that the number of covalent bonds that an atom can form is predicted from the number of electrons needed to fill its valence shell.

Introduce them to the types of covalent bonds and explain how atoms form single, double and triple bonds, using H_2, O_2 and N_2 as examples. Have the students discover the types of covalent bonds that exist in CO_2, C_2H_4 and C_2H_2.

After the students understand the three types of bonds (single, double and triple bonds) have them perform Activity 3.8. Give this activity to students either in groups or individually, depending on the number of students in the class. Next, assign the task as given in the activity in the text book. Have students show their models to the rest of the class. Finally harmonize the discussion with your conclusions.

i. Only H_2 has a single bond

ii. Both O_2 and C_2H_4 contain double bonds.

iii. N_2 contain a triple bond

Tell the students about the relative strength of single, double and triple bonds.

Polarity in covalent molecules

Activity 3.9 helps students to compare the polar and nonpolar covalent bonds. Have them explain their opinions of the formation of the covalent bonds between H_2 and HCl. Finally, harmonize their group discussions by explaining the difference between polar and nonpolar bonds. The bond in H_2 that formed between identical atoms is nonpolar, whereas the bond in HCl is polar, due to the unequal sharing of the electron pairs.

Emphasize the difference between polar and nonpolar covalent bonds. Tell them that two atoms of equal electronnegativity form a nonpolar covalent bond by sharing electrons. On the other hand, when two atoms with different electronegativity values share electrons, the covalent bond they form is a polar covalent bond. In nonpolar covalent bonds there is equal attraction for the shared pair of electrons by the nuclei of the bonding atoms. Whereas, in a polar covalent bond, the shared pairs of electrons will be attracted towards the more electronegative atom. This atom becomes partially negative (δ^-), and the less electronegative atom partially positive (δ^+).

Molecules possessing polar covalent bonds are called **dipoles**. This is because they are positive at one end and negative at the other end. Examples of polar molecules include HCl, H_2O, NH_3, HF etc.

Then continue by presenting coordinate covalent bonds (dative bonds). Have them with coordinate covalent bond forms when the electron pairs shared between atoms are supplied only by one of the bonding atoms. For two atoms to form this bond, one of them must have a lone pair of electrons and the other an unfilled or vacant shell. This bond is also called donor acceptor-bond. The atom that supplies the electron pair for sharing is the donor.

General properties of covalent compounds

Use group discussion, demonstration and question and answer methodologies for this topic.

Before completing the topic of covalent bonding, list and explain the general properties of covalent compounds. Then, let the students perform **Experiment 3.2** in groups as given in the student's textbook. Have the students record what they observe during the experiment and then report their conclusions to the class. Finally, harmonize their ideas with the following conclusions:

Experiment 3.2

Naphthalene and iodine vaporize at low temperatures. Graphite melts at a relatively higher temperature.

Substances	Melted, vaporized or nothing happened	High or low melting point
Naphthalene	Vaporizes	
Graphite	melts	high
Iodine	vaporizes	

Non-polar substances are insoluble in polar solvents like water and ethanol, whereas polar substances are soluble in polar solvents. Hexane and benzene are non-polar solvents.

Substances	Solubility			
	Water	Ethanol	Hexane	Benzene
Naphthalene	insoluble	insoluble	soluble	soluble
Graphite	insoluble	insoluble	insoluble	insoluble
Iodine	insoluble	soluble	soluble	soluble

After harmonizing the concepts developed during Activity 3.9, continue by presenting Activity 3.10.

Give this activity to students in groups. Next, assign the tasks given in the activity. Guide the students in collecting samples of covalent compounds from the school laboratory. Have the students bring some of these samples to the classroom and then describe the physical properties of the samples to the class. The properties of covalent compounds listed in the student's text book will help students to identify and characterize covalent compounds.

Assessment

Assess each student's work to determine whether or not he or she has achieved the minimum requirement level. In order to do this, you can record the performance of each student by considering how the student takes part in discussing Activity 3.7 – 3.10, involves in presenting opinions of the group, participates in performing Experiment 3.2 and makes presentation about observations of the groups, and answers questions raised during mini-lectures or stabilization. Give them Exercise 3.4, 3.5 and 3.6 as class work or homework. Check their work and have a record about their performances. From the record you have, make sure that the suggested competencies are achieved by the students.

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.

Additional Questions

1. What is the role of the central atom when drawing the Lewis structure for a molecule?
2. Why is the CF_4 molecule nonpolar, even though it contains polar bonds?

* 3. Classify the following substances as covalent compounds or ionic compounds.

a. KBr	e. NaH
b. BeO	f. SiF_4
c. O_3	g. PCl_5
d. CBr_4	h. $Mg(NO_3)_2$

4 In general, what conditions cause two atoms to combine to form:

a. a bond that is mainly covalent?
b. a bond that is mainly ionic?

* 5. Predict the type of bond that will be formed between the following atoms:

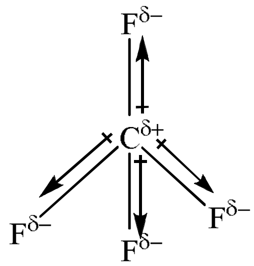
a. C and H	e. Zn and Cl
b. K and S	f. Si and O
c. H and I	
d. C and O	

Answers to Additional Questions

1. A central atom is bonded to two or more atoms. The number of lone pairs and bonding pair of electrons on the central atom of a molecule determine the geometry or structures of the molecule.
2. - In the CF_4 molecule, each C–F bond contains a polar bond, due to the electronegativity difference between C and F.

$$\overset{\delta^+}{C} - \overset{\delta^-}{F}$$

- CF_4 is a nonpolar molecule because the C–F dipole bonds cancel each other in the tetrahedral arrangements of their atoms.



Carbon tetrafluoride, CF_4

3. a. Ionic compound
   b. Covalent compound
   c. Covalent compound
   d. Covalent compound
   e. Ionic compound
   f. Covalent compound
   g. Covalent compound
   h. Ionic compound

4. The electronegativity difference between two bonded atoms is used to determine the type of bond that most likely occurs.

   i. If the electonegativity difference between the two bonded atoms is **less than 1.7**, then the bond is **mainly covalent**. For example, the bond in HCl is covalent because the electronegativity difference between H and Cl is 0.9.

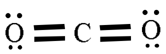
      Electronegativity of H is 2.1
      Electronegativity of Cl is 3.0 } $3.0 - 2.1 = 0.9$

   ii. If the electronegativity difference between the two bonded atoms is **greater than 1.7**, the bond is **mainly ionic**. For example, the bond in NaCl is ionic because the electronagativity difference between Na and Cl is 2.1.

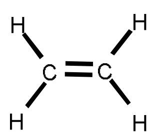
5. a. Covalent bond
   b. Ionic bond
   c. Covalent bond
   d. Covalent bond
   e. Ionic bond
   f. Covalent bond

Answers to Exercises

Exercise 3.4

1. a) CO_2  4 bonding pairs and 4 lone pairs.

   b)

   C_2H_4;  6 bonding pairs

c) N_2 ; $:N \equiv N:$ 3 bonding pairs and 2 lone pairs

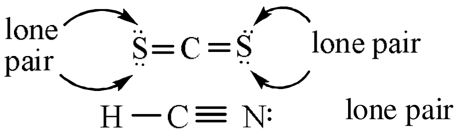
d) C_2H_2 ; $H-C \equiv C-H$ 5 bonding pairs

2.

a) CS_2 contains two double bonds, whereas HCN contains single and triple bonds.

b) $S=C=S$ for CS_2

$H-C\equiv N$ for HCN

c) 

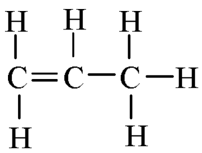
(4 lone pair)

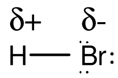
(1 lone pair)

3. Due to the presence of strong electrostatic forces between the ions, very large amount of energy is needed to overcome these strong forces. Therefore, ionic compounds have high melting points.

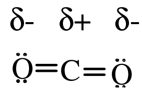
Exercise 3.5

1. a. two b. four c. six
2. a) H:H
   b) Cl—Cl
   c)



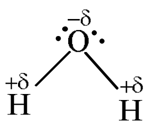
3. a. 

Polar

b. 

Nonpolar

c.



Exercise 3.6

1. ‘b’, ‘c’, ‘d’, ‘f’ and ‘i'
2. ‘a’
3. ‘a’, ‘c’ and ‘g’
4. ‘a’, ‘c’, ‘d’, ‘f’ and ‘g’

3.4 Metallic Bonds

Period Allotted 1

Competencies

After completing this section, students will be able to:

- explain the formation of metallic bond;
- explain the electrical and thermal conductivity of metals in terms of metallic bonding: make a model to demonstrate metallic bonding.

Forward Planning

Make appropriate preparation by reading about the general concepts of metallic bonding such as the formation and properties of metallic bonding. Use the student's text, reference books and other resources. Make a plan how to budget your time for the activities you will perform during the teaching-learning process and also decide when to give students the project work. You better give them before the period you intend to teach this topic.

Teaching Aid

- Model of metallic bond

Subject Matter Presentation

It is advisable to use gapped lecture, visual-based learning and collaborative learning methods to teach this topic.

Start the lesson by introducing the nature of metals. Tell the students that metals are hard and crystalline solids. Have students discuss Activity 3.11 in groups for a few minutes. Next, have them report what they did present and their ideas to the class. Finally, harmonize the discussion by showing the following figure, using a model. In the figure, the balls indicate the nuclei of metal and the empty space, a cloud of electrons.

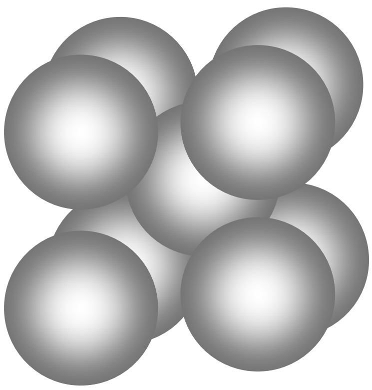


Figure 3.1 Model of a Metallic bond

Inform your students that metals can form a matrix of positively charged ions in a sea of delocalized electrons. A metallic bond is the result of electrostatic attraction between positively charged metal ions and the negatively charged delocalized electrons. The strength of a metallic bond is related to the atomic radius of the metal atom and the number of valence electrons which are delocalized.

The freedom of movement of bonding valence electrons is responsible for the high electrical and thermal conductivity that characterizes all metals. Other properties of metallic bonding contribute to unique properties of metals. For example, most metals are easy to shape, due to their malleability and ductility.

Assessment

Assess each student's work to determine whether or not he or she has achieved the minimum requirement level. In order to do this, you can use different instruments of evaluation such as class work, homework, quizzes, tests and examinations. Record the performance of each student in group discussion, presentations, and answering questions during harmonizing concepts or stabilization.

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 also arrange additional lesson time.

Additional Questions

* 1. List four physical characteristics of a solid metal.
2. In the laboratory, how could you determine whether a solid has an ionic bond or a metallic bond?

Answers to Additional Questions

1. The following are some of the typical physical properties of metals:
   - Malleability
   - Ductility
   - Conductivity
   - High melting point
   - High boiling point
2. In the laboratory, you can identify whether a solid is an ionic bond or a metallic bond, simply by knowing the fact that **metals bend when struck but ionic solids shatter.**

Answer to Exercise 3.7

1. Properties of materials are based on bonding, and the bonding in both metals and ionic compounds is based on the attraction of particles with opposite charges. However, there is a basic difference between a metallic bond and an ionic bond in terms of their bond formations, as stated below.
   - Metallic bonds are formed when metal cations attract free valence electrons. A 'sea of electrons' moves throughout the entire metallic crystal, producing a force of attraction.
   - An ionic bond is the electrostatic force of attraction between positively and negatively charged ions, and it is formed by the transfer of electrons from one atom to another.
2. Metals can conduct both heat and electricity, due to the movement of delocalized electrons in the crystal lattice.
3. No, the metallic bond is responsible only for holding metal atoms together in metallic crystals.

3.5 Intermolecular Forces

Period Allotted 4

Competencies

After completing this section, students will be able to:

- define intermolecular force;
- explain hydrogen bonding;
- explain the effects of hydrogen bond on the properties of substances;
- describe Vander Waals forces;
- explain dipole-dipole forces;
- explain dispersion forces;
- give examples of molecules with dipole-dipole forces;
- give examples of molecules in which the dispersion forces are important;
- compare and contrast the three types of intermolecular forces.

Forward Planning

Take enough time to get prepared before class by reading and by preparing the main points of your presentations on intermolecular forces such as hydrogen bonding and Vander Walls forces. Design a plan that shows the contents and activities you will treat during each period so that you can cover the entire contents of the section within four periods. Your plan should include the time allotted for the activities you perform during each period. Read the contents in the teacher's guide to get more information about the activities given in this section and the methodologies you implement.

Subject Matter Presentation

Use the group discussion and gapped lecture methodologies for this section.

You are advised to begin the lesson by discussing Activity 3.12. Have some groups present their opinions to the class. After their presentations, harmonize the discussion by concluding with the following concept:

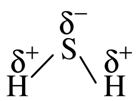
Differences in properties of ionic and covalent compounds are a result of differences in attractive forces. In a covalent compound, the covalent bond between atoms in molecules is quite strong, but the attraction between individual molecules is relatively weak. The weak forces of attraction between individual molecules are known as intermolecular forces. These forces are not strong enough to keep covalent compounds as compact solids, and hence they exist as gases or liquids.

Now continue teaching the lesson by presenting intermolecular forces.

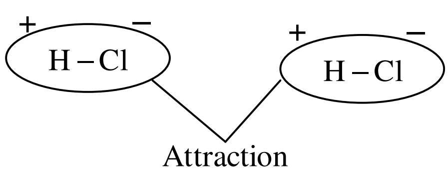
First, tell the students what intermolecular forces are and then state their types as dipole-dipole, London dispersion forces and hydrogen bonding.

Continue teaching by presenting dipole-dipole forces. Before you present the details, write the formulas HCl, HBr, H_2S and CO on the blackboard. Then, have them discuss, in groups, whether or not these molecules are polar or nonpolar for a few minutes. After that, encourage a student to present the opinions of his or her group and draw the structures of the compounds on the blackboard, showing the negative and positive ends of each molecule. Then ask him/her if many molecules of HCl or H_2S are in a container, where they attract each other or not. After the response of the student, appreciate his/her attempt and continue harmonizing his/her ideas with the facts. Tell the students that HCl, HBr, H_2S and CO are all polar molecule or dipoles. Their structures are:

$$\overset{\delta^+}{H}-\overset{\delta^-}{Cl}, \overset{\delta^+}{H}-\overset{\delta^-}{Br}, \overset{\delta^+}{C}\equiv\overset{\delta^-}{O}$$



When molecules of HCl are filled in a container, the negative end of one molecule ($Cl^{\delta-}$) attract the positive end ($H^{\delta+}$) of another molecules.



This force of attraction between the negative end of one molecule and the positive end of another molecule is a dipole-dipole force. These forces of attraction exist between molecules in polar covalent substances.

Then, continue by introducing the students to London dispersion forces. First, tell them that London forces act between all atoms and molecules. They are the only forces that exist between noble gas atoms and nonpolar molecules such H_2, O_2, Br_2 etc. Tell them also how these forces originate and why their magnitude increases with increasing atomic number or molecules mass.

After that, continue by introducing hydrogen bonding. Before presenting the details, write the following questions on the blackboard.

a) Are the molecules NH_3, H_2O and HF polar or nonpolar?

b) What type of intermolecular forces exist between molecules of NH_3, H_2O and HF?

Have the students discuss these questions in groups for a few minutes and then ask some groups to present their opinions to the class. Then harmonize the concepts suggested by the students with the truth. Tell them that the molecules H_2O, NH_3 and HF are polar. The intermolecular forces that exist in H_2O, NH_3 and HF are hydrogen bonds.

Here, emphasize that hydrogen bonding is a special type of dipole-dipole forces. But hydrogen bonding is stronger than ordinary dipole-dipole forces. For comparison, tell them that dipole-dipole forces are about 1% as strong as covalent bond whereas hydrogen bonding is about 5 – 10% as strong as a covalent bond. Next, tell that hydrogen bonding exists in polar compounds containing hydrogen atoms bonded to F, O or N. Give them additional examples of compounds containing hydrogen bonding.

Emphasize the following points to conclude the group discussions. Dipole-dipole forces are attractive forces between molecules that possess dipole moment. Dispersion forces are attractive forces that arise as a result of temporary dipoles induced in atoms or molecules. At very low temperatures (and reduced atomic speeds), dispersion forces are strong enough to hold atoms together, causing the gas to condense. The attraction between nonpolar molecules can be explained similarly.

Assessment

Assess each student's work to determine whether or not he or she has achieved the minimum required level. In order to do this, use the students' performance list and record how every student

- participates in discussing Activity 3.12 and presents the opinions of the group.

- involves to discuss the activity given in the teacher's guide and suggest opinions of the groups.
- answers questions you ask during gapped lectures, harmonizing concepts or stabilization.

Give them Exercise 3.8 as class work or homework and also a test. Check how they attempted the questions and record their achievements. Based on the record, see whether or not students have fulfilled the competencies suggested for the section.

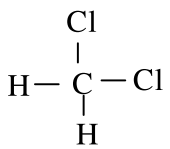
Praise students working above the minimum required level and recognize their achievements. Encourage them to continue working hard and not to become complacent. Help students working below the minimum requirement level by giving them extra activities so that they will catch up with the rest of the class. Prepare questions related to the contents in intermolecular forces and give them additional exercise or arrange extra lesson time.

Additional Questions

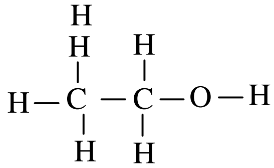
* 1. Rank the following bonds and forces from the strongest to the weakest: Covalent bond, London forces, hydrogen bonds, dipole-dipole forces.

* 2. What type of intermolecular attraction (hydrogen bonds, London forces, dipole-dipole forces) would you expect to be dominant between molecules of the following substances?

a. NH_3
b. BF_3
c. HCl
d. C_2H_6
e. H_2O_2
f.



g.



3 Use intermolecular forces to explain why oxygen is a gas at room temperature and water is a liquid.

Answers to Additional Questions

1. The decreasing order in the strength of bonds and forces is:

Covalent bonds > hydrogen bonds > dipole-dipole forces > London forces

2. a. Hydrogen bond
   b. London forces
   c. Dipole-dipole forces
   d. London forces
   e. Hydrogen bonds
   f. Dipole-dipole forces
   g. Hydrogen bonds

3. This is due to hydrogen bonds and London forces that exist in each molecule.
   - Hydrogen bonds that exist in water molecules explain why water is a liquid at room temperature.
   - The stronger London forces that exist in oxygen molecules explain why oxygen is a gas at room temperature.

Answer to Exercise 3.8

1. d 2. d

Critical thinking – This is due to the hydrogen bonding that exists in water.

Answers to the Review Questions and Problems on Unit - 3

Part I

1. A 2. B 3. C 4. C 5. C

Part II

6. metal, nonmetal
7. nonmetal
8. metals
9. chemical bonds

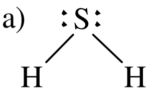
Part III

10. a) A bond formed by sharing of electrons. E.g., HCl
    b) A bond formed by electrostatic attraction between oppositely charged ions. E.g., $CaCl_2$
    c) A bond formed by a sharing of electrons in which the shared electrons are contributed by one of the atoms. E.g., NH_4^+
    d) It is a dipole-dipole interaction in which hydrogen is bonded to highly electronegative elements such as O, N, and F. e.g., H_2O
    d) Refer to the textbook, page 87.

11. a) Ionic
    b) Covalent
    c) Ionic
    e) Ionic
    f) Ionic
    g) Covalent

d) Ionic

h) Covalent

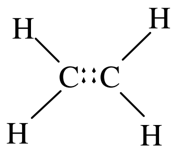
12. a) 

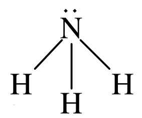
b) $Ca^{2+}[:\ddot{\underset{..}{S}}:]^{2-}$

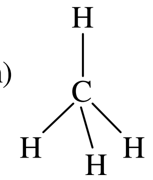
c) $2Al^{3+}3[:\ddot{\underset{..}{O}}:]^{2-}$

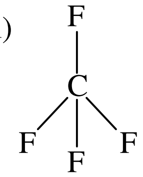
d) $H:\ddot{\underset{..}{F}}:$

e) $:N::N:$

f) 

g) 

h) 

i) 

i) $[:N \equiv O:]^{+}$

j) $Ca^{2+}2[:\ddot{\underset{..}{Cl}}:]^{-}$

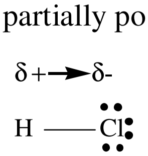
13. e, f and g

14. b and c have hydrogen bonds.

15. a) Nonpolar
    b) Polar
    c) Nonpolar
    d) Nonpolar
    e) Nonpolar
    f) Nonpolar
    g) Polar
    h) Polar

16. Any diatomic molecule that has a polar bond will be a polar molecule with its characteristic dipole moment. So, a polar molecule possesses a partial charge with a dipole moment greater than zero.

17. In an HCl molecule, the electronegativity of chlorine (3.0) is higher than that of hydrogen (2.1). Thus, the chlorine will be partially negative, and the hydrogen will be partially positive.



In a Cl_2 molecule, the two chlorine atoms have equal electronegativity values. A Cl_2 molecule has no dipole moment, and no polarity is observed.

18. **Intermolecular attractive force** is between molecules, as in two molecules of H_2O, whereas **intramolecular attractive force** is within a molecule such as CH_3COOH.

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Unit **4**

Chemical Reactions and Stoichiometry

Unit Overview — Total Period Allotted 37

This unit introduces the basic principles of chemical reactions and chemical equations. Types of chemical reactions are treated with various examples. Stoichiometric calculations, oxidation-reduction reactions, rates of chemical reactions and chemical equilibrium are presented in the unit.

Section 4.1 introduces the definition of chemical reactions and some common examples of chemical changes.

Section 4.2 presents the fundamental laws of chemical reactions and explains how they can be applied for any type of reactions.

Section 4.3 presents chemical equations and the steps necessary to write an equation. It also explains how to balance a chemical equation by using the inspection and LCM methods.

Section 4.4 presents the energy changes in chemical reactions. It defines and describes exothermic and endothermic reactions by using energy diagrams.

Section 4.5 presents all four types of chemical reactions. The categories of reaction as combinations, decomposition, single and double displacement reactions are described with various examples.

Section 4.6 presents stoichiometric calculation in chemical equations. Various examples are given for each of these stoichiometric problems.

Section 4.7 describes oxidation-reduction reactions. It defines and explains oxidation, reduction, oxidizing agent and reducing agent.

Section 4.8 presents reaction rate and how it is affected by different factors. It also explains the principle of chemical equilibrium and describes the effect of changes in temperature, pressure and concentration on the state of chemical equilibrium.

The methodologies suggested for teaching in this unit are group discussion, gapped-lecture, demonstration, experiment, answer and question and drawing a map as methodologies.

Unit Outcomes

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

- understand fundamental laws of chemical reactions and how they are applied;
- develop skills in writing and balancing chemical equations;
- understand energy changes in chemical reactions;
- know types of chemical reactions;
- develop skills in solving problems based on chemical equations (mass-mass, volume-volume and mass -volume problems);
- develop skills in determining the limiting reactant, theoretical yield, actual yield and percentage yield,
- understand oxidation-reduction reactions and analyze redox reactions by specifying the oxidizing agent and reducing agent, the substance reduced and oxidized,
- understand the rate of a chemical reaction, the state of chemical equilibrium and factors affecting equilibrium,
- demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, communicating, measuring, asking questions, designing experiments, interpreting data, drawing conclusions, applying concepts, relating cause and effect and problem-solving.

Main Contents

4.1 Introduction

4.2 Fundamental Laws of Chemical Reactions

4.3 Chemical Equations

4.4 Energy Changes in Chemical Reactions

4.5 Types of Chemical Reactions

4.6 Stiochiometry

4.7 Oxidation and Reduction Reactions

4.8 Rate of Chemical Reactions and chemical equilibrium.

4.1 Introduction

Period Allotted 1

Competencies

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

- define chemical reaction,
- give some examples of chemical reactions.

Forward Planning

Prepare yourself on the general concept of chemical reactions by reading from the student's text, reference books and other resources.

Plan how to manage students when they discuss the start-up activity, Activity 4.1 and 4.2. Your plan should include the time allotted for students to discuss the activities, present their opinions, harmonizing concepts, mini-lecture and stabilization.

Subject-Matter Presentation

You can use group discussion, question and answer as a method for teaching this topic.

You are advised to begin the class with the start up activity given in the student's textbook. The start-up activity is designed to help students understand chemical reactions and the laws governing chemical reactions. The activity helps them analyze how the reactants react to give a product. The ideas of mass-mass, volume-volume, mole-mole and limiting reactant ratios are incorporated in the activity.

Have the students discuss the start-up activity in groups. Then invite a student to present the ideas of a group to the class. Then, ask students in other groups whether the opinion of their groups is similar to those that the student presented to the class.

After the discussion, harmonize it by presenting these ideas:

i) When hydrogen burns in oxygen, the product is steam water vapor. The reaction can be classified as a combination or synthesis reaction.

ii) The balanced equation is: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

iii) Volume ratio is: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

2vol 1vol 2vol

iv) Mass ratio is: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

4 g 32 g 36 g

After harmonizing concepts related to the start-up activity, you are advised to begin the lesson with Activity 4.1. The activity is designed to help students realize that chemical reaction can serve as sources of energy. Let the students discuss this activity in groups and present their findings to the class. Finally, organize a whole-class discussion. After the discussion, harmonize it by presenting these ideas:

1. The main rocket thruster of a space shuttle is powered by a simple chemical reaction. Fuel gases are ignited in the combustion zone of the rocket and yield CO_2 and H_2O. The intense heat drives these exhaust gases downward and produces the tremendous thrust necessary to boost the rocket.
2. Many reactions provide evidence that they have occurred. Some of these indications are: changes in temperature, color, odor, gas bubbles and physical states.

Next, define a chemical reaction, reactants and products. Inform the students that all chemical reactions result in changes in the composition, properties, and energy of the original substances. After that, continue with Activity 4.2. The activity enables students discover the chemical reactions that occur in their daily lives and identify the reactants and products of these reactions. Have students discuss the activity for a few minutes in groups and then ask some groups to present their findings to the class. After the presentations, harmonize concepts:

The following are some common examples of chemical reactions.

- Burning wood
- Fermenting tella
- The heating of sugar
- The rusting of iron
- The digestion of food
- The souring of milk

Assessment

You can assess each student's work on this section by watching how the student involves in group discussions and presentation. You can also give class work by preparing questions of your own related to chemical reaction, reactants and products. Check their work and record their performances. See whether the suggested competencies are fulfilled or not. Appreciate students working above the minimum requirement level and assist those working below the minimum requirement level.

4.2 Fundamental Laws of Chemical Reactions

Period Allotted 2

Competencies

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

- state the law of conservation of mass and illustrate it, using examples;
- demonstrate the law of conservation of mass, using simple experiments,

- state the law of definite proportions and illustrate it, using examples;
- demonstrate the law of definite proportions, using a simple experiment;
- state the law of multiple proportions and illustrate it, using examples.

Forward Planning

This sub-unit presents the fundamental laws of chemical reaction.

Read the contents thoroughly familiarize yourself with the law of conservation of mass, the law of definite proportion and law of multiple proportions. Plan which contents, activities and experiments you will treat during each period. In your plan, indicate the time required for every activity you perform in each period. There are two experiments (Experiments 4.1 and 4.2) in the section. Make sure that the chemicals and apparatuses required to carry out the experiments are available in your school chemistry laboratory. Carry out the experiments yourself before you allow students to perform them.

Teaching Aids

Refer to the student's text for the chemicals and apparatus required to perform experiment 4.1 and 4.2.

Subject Matter Presentation

Use group discussion, experiment and question and answer methodologies to present the lesson.

You are advised to begin the lesson with Activity 4.3. The activity helps students discover that the law of conservation of mass always holds true for all types of reactions. Have the students perform the activity and discuss it in groups for a few minutes. Then, ask some students to present their conclusions to the class. After their presentations, hold a whole-class discussion. Following the discussion, harmonize it by presenting these ideas:

When wood burns in air (or oxygen), it is not only the ash that is formed as a product, but also the gaseous product (CO_2). The mass of the wood plus the mass of oxygen from air equals the mass of the ash plus the mass of the gaseous products. Therefore, the mass of reactants equals the mass of products. There is no loss in mass in a chemical reaction.

Then, state the law of conservation of mass. To prove this law, let the students do **Experiment 4.1** in groups as given in the student textbook. Let the students record what they observe during the experiment and report their findings to the class. Harmonize their observations with the following concepts:

Experiment 4.1

1. A white precipitate of silver chloride (AgCl) is formed. The chemical equation is:

$$\underbrace{NaCl + AgNO_3}_{\text{Before reaction } (m_1)} \longrightarrow \underbrace{AgCl + NaNO_3}_{\text{After reaction } (m_2)}$$

2. The masses m_1 and m_2 are equal ($m_1 = m_2$).
3. From the experiment, it can be concluded that mass is always conserved in the chemical reaction. Therefore, the law of conservation of mass is obeyed.

Then, introduce students to the law of a definite proportion. After that, demonstrate **Experiment 4.2** to the class. Let students write a report about their observations during the demonstration and present them to the class. Harmonize their observation with the following facts:

Experiment 4.2

1. In each case, 0.8 grams of copper (or 80% of copper) is produced.

$$CuO \Rightarrow \frac{63.5\text{g of Cu}}{79.5 \text{ g CuO}} \times 1\text{g} = 0.80\text{g Cu} \approx 80\% \text{ Cu}.$$

2. Because in each case, we reduced the same compound, copper (II) oxide (CuO) by H_2.

$$CuO + H_2 \rightarrow Cu + H_2O$$

3. No matter which method is used to prepare copper (II) oxide, it always contains 80% by mass of copper and 20% by mass of oxygen. The composition of copper (II) oxide is thus constant. Therefore, in forming CuO, copper combines with oxygen in a definite proportion.

After the demonstration, introduce students to the law of multiple proportions. Next, let them discuss Activity 4.4 in groups for a few minutes and then ask students from two different groups present their conclusions to the class. Following their presentations, harmonize concepts:

The table given below illustrates the law of multiple proportions by considering the five oxides of nitrogen. The masses of oxygen in the five compounds that combine with a fixed mass (1 gram) of nitrogen are 0.57, 1.14, 1.71, 2.28 and 2.86, respectively. Dividing each mass of oxygen by the smallest one would give the small whole numbers of 1, 2, 3, 4 and 5.

Table 4.1 Illustration of the law of multiple proportions.

Compound	Molecular formula	Mass ratio of N to O (N:O)	Ratio of oxygen /dividing each by the smallest/
Dinitrogen monoxide	N_2O	28:16 $\Rightarrow$ 1:0.57	1
Nitrogen monoxide	NO	14:16 $\Rightarrow$ 1:1.14	2
Dinitrogen trioxide	N_2O_3	28:48 $\Rightarrow$ 1:1.71	3
Nitrogen dioxide	NO_2	14:32 $\Rightarrow$ 1:2.28	4
Dinitrogen pentoxide	N_2O_5	28:80 $\Rightarrow$ 1:2.86	5

Assessment

You may assess each student's work throughout section 4.2 by supervising how he or she involves in discussing Activity 4.3 and 4.4, takes part in presentations, participates in performing Experiment 4.1 and 4.2 as well as his or her role in presenting observations of the group on the experiment to the class and then recording his/her performance.

You can also give exercise 4.1 as class work or homework. Check how every student attempted the questions and record their achievements. Ask them oral questions related to the three laws. From the record you have, make sure that the competencies suggested to the section are achieved. Praise the students working above the minimum requirement level and assist those working below the minimum requirement level.

Additional Questions

1. Classify the following as physical or chemical properties of substances.

   a. Color
   b. Digestion
   c. Respiration
   d. Ductility
   e. melting point
   f. Rusting
   g. Expansion
   h. Reactivity
   i. Solubility

2. A 15.5g sample of X combines with a 42.5g sample of Y to form the compound XY_3. What are the masses of the reactants and products? What can you conclude from the results?
3. When mercury oxide is heated, it decomposes into mercury and oxygen. If 28.4g of mercury oxide decomposes, 2.0g of oxygen is produced. What is the percent by mass of mercury in mercury oxide? Which law is obeyed?

Answer to Additional Questions

1. a. Physical property
   b. Chemical property
   c. Chemical property
   d. Physical property
   e. Physical property
   f. Chemical property
   g. Physical property
   h. Chemical property
   i. Physical property

2. According to the law of conservation of mass, the mass of reactants is equal to the mass of products in a chemical reaction.

   Therefore, the masses of reactants, X and Y = 15.5g + 42.5g = 58.0g; the mass of product, XY_3 is also 58.0g.

3. The mass of mercury is 28.4g – 2.0g = 26.4g

Therefore, $\%Hg = \dfrac{26.4g}{28.4g} \times 100 = 92.96\%$; $\%O = \dfrac{2.0g}{28.4g} \times 100 = 7.04\%$

Hence, this indicated that it obeys the law of definite proportions.

Answers to Exercise 4.1

1. a) Chemical
   b) Physical
   c) Chemical
   d) Chemical
   e) Chemical
   f) Chemical
   g) Physical
   h) Chemical

2. **Compound A**

   Fe = 1 g

   Cl = 1.27 g

   Compound B

   Fe = 1 g

   Cl = 1.9 g

   By comparing 1.27 g of Cl with 1.9 g of Cl, in the two compounds having the same mass of iron, it is possible to obtain a **simple ratio** as follows:

   $$\frac{1.27\text{g Cl in Compund A}}{1.9\text{g Cl in Compound B}} = \frac{0.668}{1} = \frac{2}{3} = 2:3$$

   Chlorine in the two compounds, A and B, is in the simple whole number ratio of 2:3, and therefore the compounds obey the law of multiple proportions.

3. NO. The mass of the substances before and after the reactions are the same. In both cases, the mass of reactants and products are equal even though the ways the reactions proceed are different.

   i) Oxygen from the air (moist air) reacts with iron to form solid iron oxide. The increase in mass is due to oxygen from the air.

   ii) During the burning of a match, carbon dioxide escapes into the air. Loss of CO_2 is the causes for the decrease in mass.

4. Dalton proposed the atomic theory which helps to explain the conservation of mass (matter) in chemical reactions as being the result of the separation, combination, or rearrangement of atoms.

4.3 Chemical Equations

Period Allotted 3

Competencies

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

- describe the conventions used to write chemical equations;

- balance chemical equations using the inspection method;
- balance chemical equations using the least common multiple (L.C.M) method.

Forward Planning

Read the contents in section 4.3 to get information about chemical equations, how to write and balance equations. Design a plan that shows which topic and activity you will treat during each period so that the whole content in the section can be covered within three periods. Your plan should include the time allotted for every activity you perform in each period. You also need to read the teacher's guide to get information about the methodologies you apply and gain more ideas about Activity 4.5 and 4.6.

Subject Matter Presentation

Writing Chemical Equation

Use group discussion and question and answer methodology to teach this topic.

You are advised to begin the lesson with Activity 4.5. The activity is designed to help student discover the differences between chemical equation and chemical reaction. Let the students perform and discuss the activity in their groups and then share their ideas with the class. After the discussion, please harmonize it by presenting these ideas:

1. **A chemical reaction** is a process in which the reacting substances, called reactants, are converted to new substances called products. In the process, a chemical change always occurs.

   A chemical equation is a shorthand expression of a chemical reaction. A chemical equation shows the rearrangement of atoms that are involved in the chemical reaction.

2. The law of conservation of mass states that mass can neither be created nor destroyed in a chemical reaction. Dalton suggested that reactions were simply rearrangements of the same numbers of atoms. A balanced chemical equation contains the same number of each kind of atom on each side of the equation. A balanced chemical equation therefore obeys the law of conservation of mass.

Next to that, tell students the steps for writing chemical equations. Give them some examples. Inform them of the conditions that any chemical equation should fulfill as well as its quantitative and qualitative meaning.

Balancing Chemical Equations

It is advisable if you use gapped lecture to teach the lesson on this topic. You can start teaching this subtopic by asking the in text question and presenting the need for balancing chemical equations. After you get responses from the students, give the answers to the questions. Then, introduce balancing chemical equations, using the inspection and LCM methods. Let students do exercises.

Assessment

You can assess each student's work throughout section 4.3 by recording how every student is performing his/her daily activity during the teaching learning process.

You can make a record in relation to:

- participation of the student in group discussion.
- participation of the student in presentation.
- answering questions during gapped lectures.

You may give them Exercises 4.2 and 4.3 as class work or homework and quiz or test. Check their work and record their achievements. Based on the information you get from the record in students performance list, check whether or not the suggested competencies to section 4.3 are fulfilled.

Appreciate students working above the minimum requirement level and recognize their achievements. For students working below the minimum requirement level, give the necessary assistance in class or arrange additional lesson time. You can give them the questions recommended to them from the additional questions of this section or from your own so that they will catch up with the rest of the class.

Additional Questions

1. When gasoline is burned in an automobile engine, what evidence indicates that a chemical change has occurred?
2. Why is it important to reduce coefficients in a balanced equation to the lowest possible whole-number ratio?

* 3. After balancing a chemical equation, what must be the same on both sides of the equation?

* 4. Write the balanced chemical equations for each of the following reactions.
   a. Nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.
   b. Butane gas (C_4H_{10}) burns in air, producing carbon dioxide and water vapor.
   c. Iron (III) chloride reacts with sodium hydroxide, producing iron (III) hydroxide and sodium chloride.
   d. Heating sodium hydrogen carbonate produces sodium carbonate, carbon dioxide and water.

5 Change the following chemical equations to word equations.
   a. $PCl_3 + Cl_2 \rightarrow PCl_5$
   b. $KBr + AgNO_3 \rightarrow AgBr + KNO_3$

Answers to Additional Questions

1. In automobile engines, gasoline is combined with oxygen, producing *carbon dioxide*, *water* and *energy* that powers the automobile.
2. The smallest whole-number coefficients in a balanced chemical equation tell us the smallest number of substances that are involved in the reaction.
3. The number of atoms of each element in a balanced chemical equation must be equal on both sides of the equation.
4. a. $4NO_2 + O_2 \rightarrow 2N_2O_5$
   b. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$
   c. $FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$
   d. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
5. a. Phosphorous trichloride + Chlorine → Phosphorous pentachloride
   b. Potassium bromide + Silver nitrate → Silver bromide + Potassium nitrate

Answers to Exercises

Exercise 4.2

1. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$
2. $CaCO_3 \longrightarrow CaO + CO_2$
3. $2H_2O_2 \longrightarrow 2H_2O + O_2$
4. $2Al + 2H_3PO_4 \longrightarrow 2AlPO_4 + 3H_2$
5. $2HNO_3 + 3H_2S \longrightarrow 2NO + 3S + 4H_2O$

Exercise 4.3

1. a) $2SO_2 + O_2 \longrightarrow 2SO_3$
   b) $2KClO_3 \longrightarrow 2KCl + 3O_2$
   c) $Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$
   d) $2AgO \longrightarrow 2Ag + O_2$
2. a) $PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$
   b) $Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2$

c) $2Zn(NO_3)_2 \xrightarrow{\Delta} 2ZnO + 4NO_2 + O_2$

d) $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$

f) $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$

4.4 Energy Changes in Chemical Reactions

Period Allotted 3

Competencies

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

- explain energy changes in chemical reactions;
- define endothermic and exothermic reactions;
- describe endothermic and exothermic reactions;
- illustrate endothermic and exothermic reactions, using diagrams;
- conduct simple experiments that demonstrate exothermic and endothermic reactions;
- describe the importance of chemical changes in the production of new substances and energy.

Forward Planning

Prepare yourself by reading about energy changes in chemical reactions. This sub-unit includes Experiments 4.3 and 4.4. Try them yourself before the class. Check whether or not the materials and chemicals are available in the laboratory for experiment 4.3 and experiment 4.4 given in this sub-unit. Make a plan of your own that shows which topics, activities and experiments you will treat during each period so that you can cover the entire contents in the section within three periods. Your plan should clearly indicate how to budget your time for every activity you perform in each period.

This sub-unit requires a teaching aid. Prepare a big chart or diagrams that illustrate endothermic and exothermic reactions. Use a sturdy material – for example, cardboard for your chart.

Teaching Aids

- Diagrams that illustrate endothermic and exothermic reactions
- Refer to the student's text for the chemicals and apparatus required to perform Experiments 4.3 and 4.4.

Subject Matter Presentation

Use group discussion demonstration, and question and answer methodologies to teach this section.

Endothermic and Exothermic Reactions

You are advised to begin the lesson with Activity 4.6. Let the students discuss the activity in groups and then have them share their ideas with the class. Next, hold a whole-class discussion. After that, harmonize their ideas with the facts as follows:

When you bake bread, the bread absorbs energy by absorbing heat. Some of the heat is used to remove the water from the dough and the rest of the heat is used to heat and dry the bread.

Next, explain energy changes in chemical reactions, using examples. Define exothermic and endothermic reactions. Bring a big chart or diagrams to the class (that you already prepared as a teaching aid) to illustrate endothermic and exothermic reactions. Let them know the energy diagrams of endothermic and exothermic reactions. Inform them about the significance of negative and positive values of enthalpy changes.

Let students do **Experiment 4.3** in groups, as given in the student textbook. Have the students write down what they observe in the experiment and present their observations to the class.

Experiment 4.3

1. The outside of the beaker feels hot due to the release of heat from the system to the surroundings.
2. Steam is observed coming out.
3. Black.
4. $C_{11}H_{22}O_{11} + H_2SO_4 \rightarrow C + 12H_2SO_4$
5. Since heat is given out, the reaction is exothermic.

Let students do **Experiments 4.4** in groups, as given in the student textbook. Have the students write down what they observe in the experiment and present their observations to the class.

Experiment 4.4

1. Cold
2. Decreased
3. Absorption of heat from the surroundings to the system. A fall in the reading of the thermometer indicates that the reaction is endothermic.

Importance of Chemical Changes

You are advised to begin the lesson with Activity 4.7. Let the students discuss the activity in groups for a few minutes and then ask them to share their ideas with the class. Next, hold a whole-class discussion. After that, harmonize the suggested ideas with the truth. The important functions of chemical changes may include producing energy as heat, light or electricity or the energy required to move motor vehicles and even to provide our body with energy.

Assessment

Assess how every student is working throughout section 4.4. You can do this by recording the performance of each student. To make records, see how each student:

- involves in discussing Activity 4.6 and 4.7.
- takes part in presenting opinions of the group after discussion.
- answers questions raised during mini-lectures.
- participates in performing Experiment 4.3 and 4.4.
- takes part in presentation about the observations of the group during experiments.

You can also give Exercise 4.4 and other questions of your own as class work or homework. Correct their work and record their achievements.

Based on the record you have, evaluate whether or not the students have achieved the competencies suggested to the section. Appreciate students working above the minimum requirement level and give them extra work. Assist those working below the minimum requirement level by arranging additional lesson time or giving them exercises on points they didn't understand. You can use the additional questions suggested to them and some other questions from reference materials.

Additional Questions

* 1. Why does ΔH for an exothermic reaction have a negative value?

* 2. Classify the following reactions as exothermic or endothermic. In each case, determine the sign of ΔH and draw an enthalpy diagram.

a. $4Al + 3O_2 \rightarrow 2Al_2O_3 + 3352\ kJ$

b. $CO_2 + 2H_2O + 891\ kJ \rightarrow CH_4 + 2O_2$

c. $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2 + \text{energy}$

d. $CaCO_3 + \text{energy} \rightarrow CaO + CO_2$

Answers to Additional Questions

1. $\boxed{\Delta H = H_{product} - H_{reactant}}$

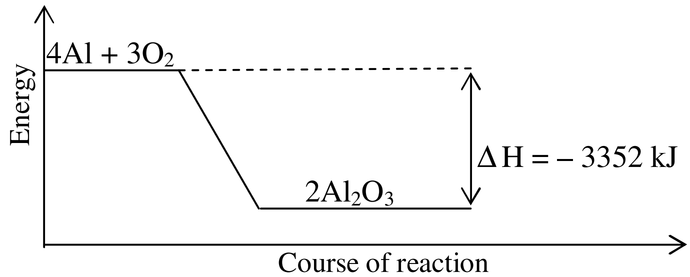
   For exothermic reaction, $H_{product} < H_{reactant}$.

   When $H_{reactant}$ is subtracted from the smaller $H_{product}$, a negative value for ΔH is obtained.

2. a. $4Al + 3O_2 \rightarrow 2Al_2O_3 + 3352$ kJ

   The reaction is exothermic $\Rightarrow$ $\Delta H =$ negative (–3352kJ).

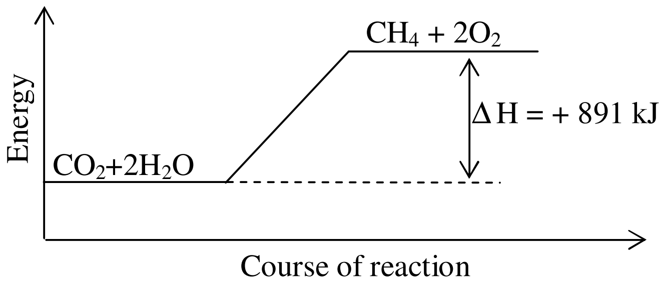
   The following figure is the energy diagram for the reaction.



   b. $CO_2 + 2H_2O + 891$ kJ $\rightarrow CH_4 + 2O_2$

   The reaction is endothermic $\Rightarrow$ $\Delta H =$ positive (+891 kJ)

   The following figure illustrates the energy diagram for the reaction.

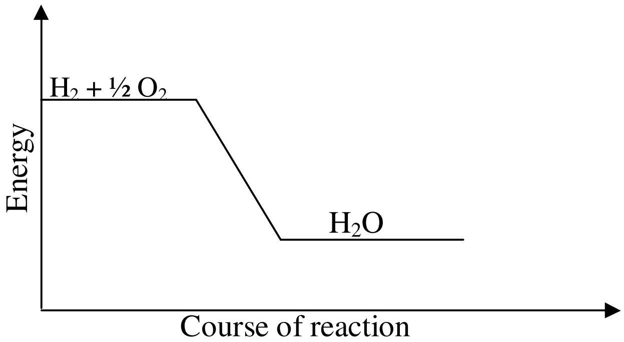


   c. Exothermic reaction

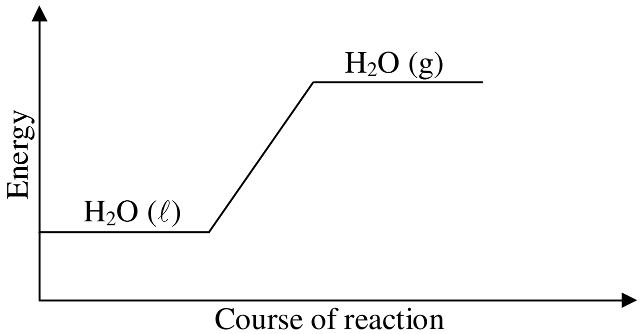
   d. Endothermic reaction

Answer to Exercise 4.4

a) Exothermic reaction; $\Delta H =$ negative (–285.8 kJ)



b) Endothermic reaction; ΔH = positive (+40.7 kJ)



4.5 Types of Chemical Reactions

Periods Allotted 3

Competencies

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

- list the four types of chemical reactions;
- define combination reaction and give examples;
- conduct some experiments on combination reactions;
- define decomposition reaction and give examples;
- conduct some experiments on decomposition reactions;
- define single displacement reaction and give examples;
- conduct some experiments on single displacement reactions;
- define double displacement reaction and give examples;
- conduct some experiments on double displacement reactions.

Forward Planning

Read the contents on the types of chemical reactions thoroughly. Prepare a plan in relation to the contents, activities and experiments you are going to deal with during each period so that all contents in the section will be covered within three periods. Set questions on points which you intend to teach during each period and use them when you are harmonizing concepts or during stabilization. Also budget your time for every activity you perform during each period.

There are four experiments in this section (Experiments 4.5, 4.6, 4.7 and 4.8). Prepare the chemicals and apparatuses required to perform the experiments. Carry out the experiment yourself before you allow students to perform them.

Teaching Aids

- Refer to the students' text for the chemicals and apparatus required to perform Experiments 4.5, 4.6, 4.7 and 4.8.

Subject Matter Presentation

Use the group discussion, experiment, demonstration and question and answer methods to teach this section.

You are advised to start the lesson with Activity 4.8. Let the students discuss it in groups and then present their findings to the class. After the discussion, harmonize it by presenting these ideas:

a. In our body, we host a type of chemical reaction known as **combustion**. In this reaction, carbohydrates react with inhaled oxygen and produce energy, carbon dioxide and water.

b. During the fermentation of 'tej' or 'tella', the yeast ferments glucose to produce the alcohol ethanol. The general chemical reaction can be given as:

$$C_6H_{12}O_6 \rightarrow 2CH_3CH_2OH + 2CO_2$$

c. The complete combustion of kerosene, like that of other hydrocarbons, gives heat energy, carbon dioxide and water.

After harmonizing concepts related to the activity, list the types of chemical reactions and continue presenting combination reaction.

Demonstrate experiment 4.5 to the class as given in the student's textbook. Let the students write down their observations and report them to the class. Next, harmonize concepts using the following points:

Experiment 4.5

1. Grayish and yellow colors, respectively.

2. A grayish black compound, iron sulphide, is formed in the test tube.
3. Fe + S → FeS
4. Combination reaction.

After the experiment, give them detailed information about combination reaction. Include examples. Then continue by presenting decomposition reaction. First, let students do experiment 4.6. Have the students write a report and present their observation to the class. Finally, harmonize concepts using the following points:

Experiment 4.6

1. Green color.
2. A green powder of $CuCO_3$ is decomposed to give a black residue of copper oxide (CuO), which is formed in the test tube.
3. The limewater turned milky. This indicates that carbon dioxide is given off in the reaction.
4. $CuCO_3 \xrightarrow{\Delta} CuO + CO_2$

After the experiment, continue your discussion by presenting the decomposition of nitrates and carbonates.

Then introduce single displacement reaction. First, let students do experiment 4.7 and write a report and present their observation to the class. Finally, harmonize concepts, using the following points:

Experiment 4.7

1. Copper metal is deposited on the surface of the iron plate after one day.
2. $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$
3. The more active metal, iron, displaced the less active metal, copper, from $CuSO_4$ solution. Therefore, the reaction is said to be single displacement reaction.

After the experiment, give them detailed information about single displacement reaction with various examples.

Finally, proceed to the double displacement reaction by letting students do experiment 4.8. Have them write a report and present their observations to the class. Finally, harmonize concepts using the following points:

Experiment 4.8

1. Barium sulfate precipitate is formed, and sodium nitrate remains in the solution.
2. A white precipitate.

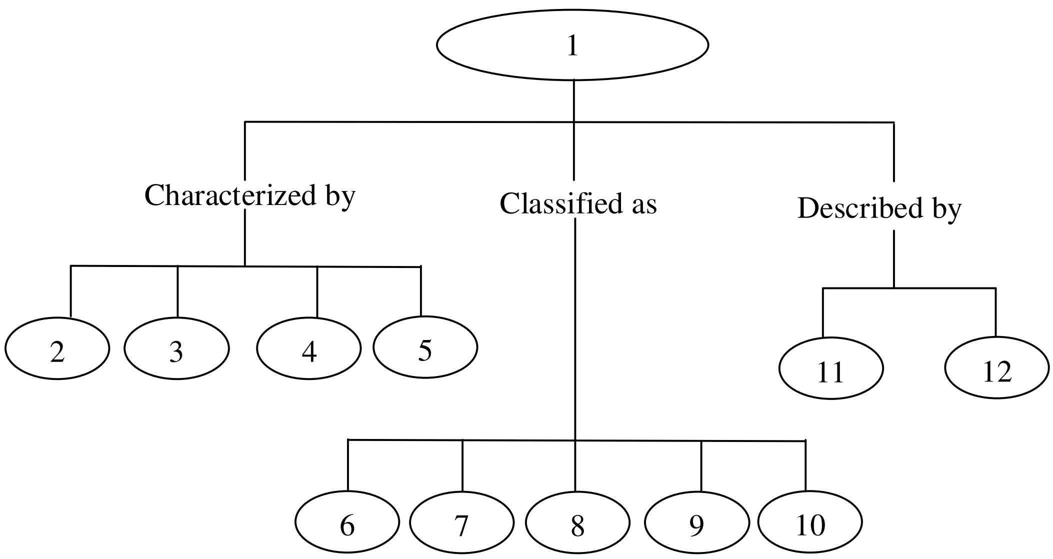
3. $Ba(NO_3)_2 + Na_2SO_4 \rightarrow \underset{\text{White}}{BaSO_4} \downarrow + 2NaNO_3$

After the experiment, continue presenting double displacement reaction using various examples. Also let the students do exercises and check whether or not they are able to identify the types of chemical reactions.

Concept Mapping

After you have presented chemical reaction, list the following terms on the blackboard and let some volunteer students try to complete the concept map using the terms:

- synthesis
- single displacement
- change in energy
- change in temperature
- word equation
- decomposition
- double displacement
- change in color
- chemical equation
- change in odor
- combustion
- chemical reaction



Appreciate their attempts in completing the concept map and then harmonize using the appropriate terms as follows:

1. Chemical reaction
2. Change in energy
3. Change in color
4. Change in odor
5. Change in temperature
6. Synthesis
7. Single displacement
8. Decomposition
9. Double displacement
10. Combustion
11. Word equation
12. Chemical equation

Assessment

Assess each student's work throughout section 4.5. Record how every student is doing in your student's performance list. You can make a record on how every student:

- participates in discussions (Activity 4.8).
- takes part in presenting concepts after discussion.
- involves in performing experiment (4.5 – 4.8).
- participates in presenting results and conclusions from the experiment.

Give them Exercises 4.4, 4.5, 4.6 and 4.7 as class work or homework. You can also give them test or quiz. Check the work of all students and record their achievements.

From your record on what students have done, see how many of them achieved the suggested competencies for section 4.5. Encourage students working above the minimum requirement level to continue working hard and also give them additional work. Assist students working below the minimum requirement level either by arranging additional lesson time or giving them additional exercise.

Additional Questions

* 1. Classify the following reactions as combination, decomposition, single or double displacement reactions.

a. $Cl_2 + 2KI \rightarrow 2KCl + I_2$

b. $H_3PO_4 + 3LiOH \rightarrow Li_3PO_4 + 3H_2O$

c. $3HNO_2 \rightarrow HNO_3 + H_2O + 2NO$

d. $CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

e. $4Fe (s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

* 2. Complete and balance the following equations.

a. $Al + CuSO_4 \rightarrow$

b. $AgNO_3 + Na_2CO_3 \rightarrow$

c. $K_2CO_3 + H_3PO_4 \rightarrow$

d. $F_2 + FeI_2 \rightarrow$

e. $BaCO_3 \rightarrow$

Answers to Additional Questions

1. a. Single displacement reaction.

b. Double displacement reaction.
c. Decomposition reaction.
d. Double displacement reaction.
e. Combination reaction.

2. a. $2Al + 3CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$
b. $2AgNO_3 + Na_2CO_3 \rightarrow 2NaNO_3 + Ag_2CO_3$
c. $3K_2CO_3 + 2H_3PO_4 \rightarrow 2K_3PO_4 + 3H_2O + 3CO_2$
d. $F_2 + FeI_2 \rightarrow FeF_2 + I_2$
e. $BaCO_3 \rightarrow BaO + CO_2$

Answers to Exercises

Exercise 4.5

1. $2CO + O_2 \rightarrow 2CO_2$
2. $H_2O + SO_2 \rightarrow H_2SO_3$
3. $Na_2O + H_2O \rightarrow NaOH$
4. $CaO + CO_2 \rightarrow CaCO_3$

Exercise 4.6

1. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
2. $Cu + Zn(NO_3)_2 \rightarrow$ No reaction
3. $F_2 + CaCl_2 \rightarrow CaF_2 + Cl_2$
4. $Br_2 + NaCl \rightarrow$ No reaction

Exercise 4.7

1. a) Single displacement reaction
b) Combination reaction
c) Double displacement reaction
d) Decomposition reaction

2. a) Single displacement reaction
b) Combination reaction
c) Double displacement reaction
d) Decomposition reaction
e) Double displacement reaction
f) Decomposition reaction

3. a) $3Mg + N_2 \longrightarrow Mg_3N_2$
b) $Na_2CO_3 \xrightarrow{\Delta}$ No reaction
c) $BaCO_3 + 2HNO_3 \longrightarrow Ba(NO_3)_2 + CO_2 + H_2O$
d) $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$
e) $FeCO_3 \xrightarrow{\Delta} FeO + CO_2$
f) $H_2CO_3 + 2NaOH \longrightarrow Na_2CO_3 + 2H_2O$

4.6 Stoichiometry

Period Allotted 10

Competencies

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

- deduce mole ratios from balanced chemical equations;
- solve mass-mass problems, based on given chemical equations;
- define molar volume;
- state Avogadro's principle;
- solve volume-volume problems, based on given chemical equations;
- solve mass-volume problems, based on given chemical equations;
- define limiting and excess reactants;
- determine limiting and excess reactants of a given chemical reaction;
- show that the amount of product of a chemical reaction is based on the limiting reactant;
- define the terms: theoretical yield, actual yield and percentage yield;
- calculate the percentage yield of a chemical reaction from given information.

Forward Planning

Read the contents of Section 4.6 (stoichiometry) thoroughly in order to get a clear information about the details. Prepare your own plan that shows the topics you are going to deal with during each period so that the entire contents of the section can be covered within ten periods. Your plan should indicate the time allotted for every activity you perform in each period.

Subject Matter Presentation

You are advised to use group discussion and gapped lecture to teach this section.

After introducing the topic, you are advised to begin the class with Activity 4.9. Let the students discuss it in groups for a few minutes and then ask them to present their findings to the class. Hold a whole-class discussion. After the discussion, harmonize it by presenting these ideas:

To assemble a complete bicycle, the mechanic should take 1 frame (body part) and 2 wheels (or 1 pair of wheel) for each complete bicycle. Based on this, 10 frames of bicycles need exactly 20 wheels (or 10 pairs of wheels). But, in the shop there are only 16 wheels (or 8 pair of wheels). Therefore, the mechanic can assemble 8 frames with 8 pairs of wheels to produce 8 complete bicycles.

- The remaining 2 frames of the bicycle are leftover in other words, they are excess.
- This shows the total number of bicycles that can be completely assembled is limited by the number of wheels. Be sure that the students relate this activity with chemical reactions.

After harmonizing the concepts, give present the definition and principles of stoichiometry. Introduce the two principles on which stoichiometric calculations are based. Inform them of the types of stoichiometric problems and then continue by presenting with mass-mass relationships.

Mass-Mass Relationships

Tell the students when to use this relationship. Explain that solving such problems can be achieved by using either the mass-ratio or mole-ratio method. Then, introduce the steps to be followed in solving problems, using the mass-ratio method ,with examples. Give them exercises and have them practice solving problems. Check how students are doing.

Next, proceed to the mole-ratio method. Introduce them to the steps with examples, and have them practice applying this method to solve problems. Give them exercise 4.8 as class work or homework and check their work.

Volume- Volume Relationships

Before you start solving problems, introduce the students to molar volume, its numerical value, STP and Avogadro's law. Emphasize that one mole of a gas can occupy a volume of 22.4 liters only at 0ºC and 1 atm. Tell them also that the converse of Avogadro's law holds true if the temperature and pressure of the gases are equal. Then, solve some problems by following appropriate steps. After that, let students practice solving volume-volume problems. You can use exercise 4.9 for this purpose. Check the work of the students and give corrections.

Mass - Volume Relationships

Make sure that students can solve volume-volume problems. Then explain how to tackle mass-volume problems. Tell them how to solve mass-volume relationships and give them some examples. After that, let students practice solving mass-volume problems. You can use exercise 4.10 for this purpose. Check how the students are doing in solving the questions and give them corrections.

Limiting and Excess Reactants

In most cases, students have difficulty in understanding limiting and excess reactants. So, you better use analogy to help them realize the concepts.

Assume that a school has 280 chemistry text books for Grade 9 students in its store. The number of students in this grade is 315. If the school distributes one book to one student, 35 students will not get books. So there are 35 students in excess. The number of student who has got books is limited to the number of books in the store. The same analogy applies to chemical reactions. If two reactants are not mixed according to their stoichiometric proportions and allowed to react, the amount of product that can be obtained from the reaction is determined by the reactant existing in smaller proportion.

Give them more examples and define limiting and excess reactants. Then, solve problems of identifying limiting and excess reactants and of determining the quantity of products that result from a given chemical reaction. Let students practice identifying the limiting reactants, excess reactants and calculating the amount of products. Give them Exercise 4.11 as class work or homework. Check their work and give corrections.

Theoretical, Actual and Percentage Yield

Before you start solving percentage yields of reactions, first define theoretical yield, actual yield and percentage yield of a reaction. Theoretical yield is the maximum amount of product we expect from a reaction if the limiting reactant is completely used up, according to the balanced chemical equation. Actual yield is the amount of product practically obtained from the reaction. After that, let them know that the actual yield is smaller than the theoretical yield. Introduce them to calculating percentage yield. Solve some problems to show how they can calculate either theoretical yield or actual yield or percent yield from given information. Finally, give them exercise 4.12 as class work or homework and check their work and give corrections.

Assessment

Assess each student's work throughout section 4.6. To do this, use students' performance list and record every performance of the student. Your records may be based on the student participation in discussing activities, in presentation, and in answering questions raised during gapped lectures.

You can also give Exercises 4.8, 4.9, 4.10, 4.11 and 4.12 as class work or homework and a test. Check the work of the students and record their achievements. Based on what you see from the performance list, make sure that most of the students have accomplished the competencies suggested for section 4.6. Appreciate students working above the minimum requirement level. Give assistance to students working below the minimum requirement level by arranging extra lesson time or giving them more exercise on the topics in which they face difficulties.

Additional Questions

1. Why do you think a balanced chemical equation is needed in solving stoichiometric calculations?

∗ 2. What mass of CH_4 formed when 8.3 moles of CO react with hydrogen according to following equation?

$$2CO + 2H_2 \rightarrow CH_4 + CO_2$$

∗ 3. Calculate the mass of calcium carbide that is needed to produce $100cm^3$ of acetylene according to this equation.

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

4. What volume of oxygen reacts with carbon monoxide to produce 0.5 mole of carbon dioxide at STP?

∗ 5. If 46.0 g of CH_4 reacts with 32.0g of O_2, according to the following reaction,

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

   a. which substance is the limiting reactant?
   b. how many grams of it remain unreacted?
   c. how many grams of CO_2 would be produced?

6. If 3 moles of calcium are reacted with 3 moles of oxygen, then
   a. which substance is the limiting reactant?
   b. how many moles of CaO are formed?

∗ 7. What is the percentage yield of a reaction in which 200g PCl_3 reacts with excess water to form 128g HCl according to the following reaction?

$$PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3HCl$$

8. When 56.6g of calcium and 30.5g of nitrogen gas undergo a reaction that has a 90% yield, what mass of calcium nitride is formed?

Answers to Additional Questions

1. Because it is used to determine the mole ratio of reactants and products.
2. 66.4 g
3. 0.29 g
4. 5.6 L
5. a. Oxygen b. 38g c. 22g
6. a. Calcium b. 3 moles
7. 80.4%
8. 62.9 g

Answers to Exercises

Exercise 4.8

1. 20.8 g

4. 2.25 mol

2. 40.8 g
3. 555 g
5. 0.033 mol
6. 5 mol

Exercise 4.9

1. 16.8 L
2. 4.8 L
3. 116 L
4. 168 L
5. 2.23 mol

Exercise 4.10

1. 64.4 L
2. 72 g
3. 0.29 g
4. 466 ml

Exercise 4.11

1. a) HCl
   b) 7.1 g
   c) 1.8 g
2. 71 g
3. 8.8 g
4. a) Ca
   b) 3 moles
5. 2 g

Exercise 4.12

1. 80%
2. Actual yield = 18.13 g. Percentage yield = 66.2%.
3. 134.25 g

4.7 Oxidation and Reduction Reactions

Period Allotted 5

Competencies

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

- define redox reactions;
- define the terms oxidation and reduction, in terms of electron transfer;
- define oxidation number (oxidation state);
- state oxidation number rules;
- determine the oxidation number of an element in a given formula;
- describe oxidizing and reducing agents;

- analyze a given redox reaction by specifying the substance reduced, the substance oxidized, and the oxidizing and reducing agents;
- distinguish between redox and non-redox reactions.

Forward Planning

Read thoroughly the contents on oxidation and reduction reactions. Design a plan of your own that shows the details of topics and activities you are going to treat during each period in such a way that the entire content will be covered within five periods. In your plan, indicate the duration of time you allot for group discussion, gapped lectures and other activities you perform during each period. Also read the teacher's guide to get information about the methodologies suggested to teach the sub-unit and to gain more ideas about Activities 4.10 and 4.11.

Subject Matter Presentation

Oxidation and Reduction

It is advisable to use group discussion and gapped lecture for this topic.

You are advised to start the lesson with Activity 4.10, which is given in the student textbook. Let your students perform and discuss the activity in their groups. Next, invite some students to present their findings to the class. Then, hold a whole-class discussion and inform them of the following points:

- The towel, the drying agent, gets **wet.** The dish, the wetting agent, gets **dry**.
    - The towel is said to be a drying agent because it dries the dish but becomes wet itself after drying the dishes.
    - The dish is said to be a wetting agent because it causes the towel to get wet, but the dish itself becomes dry.
- After a thorough discussion on this activity, let the students relate this activity with oxidation, reduction, oxidizing agent and reducing agent.

After harmonizing concepts in activity 4.10, continue by presenting oxidation and reduction. Let students suggest the definitions first. Following their responses, define the terms yourself. Inform them that oxidation and reduction are always inseparable and occur simultaneously in chemical reactions. Next, define oxidation number and introduce the rules for assigning oxidation number. Then, show how they can assign or determine oxidation number to different species. Give them exercises and let them practice assigning oxidation numbers. Check their work and give corrections.

Oxidizing and Reducing Agents

Start teaching this part by asking students to define oxidizing and reducing agents. After their responses, define the terms yourself. Then inform them that oxidizing and reducing agents can be sometimes identified by the visible color changes they show when mixed with substances that are easily oxidized and reduced, respectively. Also, explain some of the factors on which the reducing or oxidizing abilities of substances depend.

Analyzing Redox Reactions

Start the topic with Activity 4.11. Help students to form groups and let them discuss the activity for a few minutes. Then have them present the ideas from their discussions. Then, harmonize the concepts suggested by the students with the facts:

After that, ask the students about oxidation, reduction, oxidizing agents, reducing agents and oxidation number. After their responses, explain these concepts to them. Tell them that oxidation is a process by which atoms or ions lose electrons, and that reduction is a process whereby atoms or ions gain electrons.
A reducing agent supplies electrons to the substance getting reduced, and is itself oxidized because it loses electrons. Every redox reaction involves a transfer of electrons. The reducing agent (itself oxidized) supplies the electrons, and the oxidizing agent (itself reduced) accepts the electrons. Explain the differences between redox and nonredox reactions and the methods to identify these reactions.

Assessment

You can assess each student's work throughout section 4.7 by supervising how every student is doing during each period. You need to follow-up strictly the performance of every student and register what you see in the students' performance list. Recording the performances may be based on how every student:

- involves in group discussions (Activity 4.10 and 4.11)
- participates in presenting ideas of the group to the class.
- participates in answering questions raised during gapped lectures.
- answers questions given to them as activity between the gapped lectures.

Give exercises 4.13 and 4.14 as homework. You can also give a quiz or test. Check their work and record their achievements. Encourage students working above the minimum requirement level and also give them additional activity. In relation to students working below the minimum requirement level, arrange extra lesson time or give them additional exercise. The questions you give them may emphasize on definition of redox reaction, assigning oxidation number, identifying redox and non-redox reactions or identify species oxidized or reduced and oxidizing and reducing agents in a redox reaction.

Additional Questions

*1. Explain the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?

2. Explain how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?

*3. Is it possible to have a reaction in which oxidation occurs and reduction does not? Explain.

4. 4. Determine the oxidation number of each atom in the following substances.

a. Mg_3N_2 e. HPO_3^{-2}

b. P_4O_6 f. K_2O_2

c. SbF_6^- g. $K_4[Fe(CN)_6]$

d. $Ca(HSO_4)_2$

*5. Arrange the following species in order of increasing oxidation number of the sulphur atom:

a $S_2O_3^{-2}$ b. S_8 c. S^{-2} d. $S_4O_6^{-2}$ e. H_2SO_4 f. SO_2

6. Identify the oxidizing agent and the reducing agent in each of the following reactions.

a. $3Cu + 8H^+ + 2NO_3^- \rightarrow 3Cu^{+2} + 4H_2O + 2NO$

b. $MnO_2 + ClO_3^- \rightarrow MnO_4^- + Cl^-$

c. $H_2O_2 + H_2SO_4 + FeSO_4 \rightarrow Fe_2(SO_4)_3 + H_2O$

Answers to Additional Questions

1. A reducing agent supplies electrons to the substance getting reduced and an oxidizing agent accepts electrons from the other substance in a redox reaction.
2. When an atom is oxidized, its oxidation number increases, and when an atom is reduced, its oxidation number decreases.
3. Oxidation and reduction are complementary processes. Oxidation cannot occur unless reduction also occurs.
4. a. Mg = +2, N = –3

b. P = +3, O = –2

c. Sb = +5, F = –1

d. Ca = +2, H = +1, S = +6, O = –2

e. H = +1, P = +3, O = –2

f. K = +1, O = –1

g. K = +1, Fe = +2, C = +2, N = –3

5. $\overset{-2}{S}, S_8^0, \left(\overset{+2}{S_2}O_3\right)^{-2}, \left(\overset{2.5}{S_4}O_6\right)^{-2}, \overset{+4}{S}O_2, H_2\overset{+6}{S}O_4$

6. a. $3Cu^o + 8H^+ + 2\overset{+5}{N}O_3^- \rightarrow 3\overset{+2}{Cu} + 4H_2O + 2\overset{+2}{N}O$

Oxidizing agent: $\overset{+5}{N}O_3^- \rightarrow \overset{+2}{N}O$

Reducing agent: $Cu^o \rightarrow \overset{+2}{Cu}$

b. $\overset{+4}{Mn}O_2 + \overset{+5}{Cl}O_3^- \rightarrow \overset{+7}{Mn}O_4^- + Cl^-$

Reducing agent: $\overset{+4}{Mn}O_2 \rightarrow \overset{+7}{Mn}O_4^-$

Oxidizing agent: $\overset{+5}{Cl}O_3^- \rightarrow Cl^-$

c. $H_2\overset{-1}{O}_2 + H_2SO_4 + \overset{+2}{Fe}SO_4 \rightarrow \overset{+3}{Fe_2}(SO_4)_3 + H_2\overset{-2}{O}$

Reducing agent: $\overset{+2}{Fe}SO_4 \rightarrow \overset{+3}{Fe_2}(SO_4)_3$

Oxidizing agent: $H_2\overset{-1}{O}_2 \rightarrow H_2\overset{-2}{O}$

Answers to Exercises

Exercise 4.13

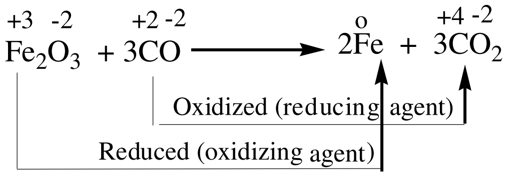
1. a) +3 e) −1
   b) −3 f) −3
   c) +2.5 g) +2
   d) +5

2. a) −2 e) +6
   b) +5 f) +7
   c) −3
   d) +5

3. a) Reduction d) Oxidation
   b) Oxidation e) Oxidation
   c) Reduction f) Reduction

Exercise 4.14

a)

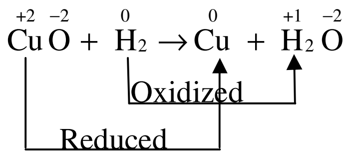


Fe is reduced; Fe_2O_3 is an oxidizing agent.

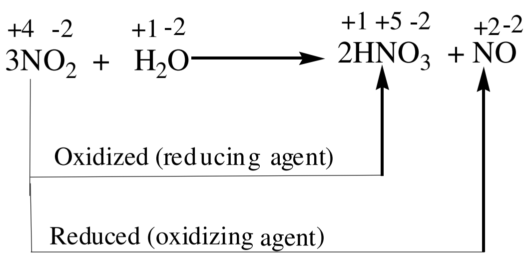
C is oxidized; CO is a reducing agent.

b) H_2 is oxidized; it is the reducing agent

Cu is reduced. CuO is an oxidizing agent.



c)



In this reaction, N undergoes both oxidation and reduction reaction. NO_2 is both a reducing agent and an oxidizing agent.

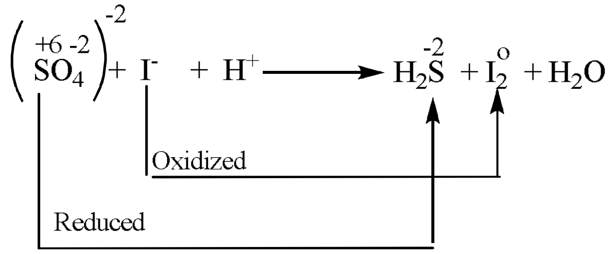
d) $\overset{+1\ +5\ -2}{AgNO_3} + \overset{+1\ -1}{NaCl} \longrightarrow \overset{+1\ -1}{AgCl} + \overset{+1\ +5\ -2}{NaNO_3}$

This is not a redox reaction because there is no change in the oxidation numbers of the substances.

e) Only oxidation (reducing agent)

f) Copper is oxidized, and Nitrogen in HNO_3 is reduced.

g)



h) Mn in MnO_4^- is reduced; MnO_4^- is an oxidizing agent.

2. a) Reduction
   b) Reduction
   c) Oxidation
   d) Reduction
   e) No change

4.8 Rate of Chemical Reaction and Chemical Equilibrium

Period Allotted 10

Competencies

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

- define reaction rate;
- describe reaction rate using graphs;
- carry out an experiment to illustrate the relative rate of reactions;
- list the preconditions for a chemical reaction to occur;
- explain how collision, activation energy and the proper orientation of reactants cause a chemical reaction to occur;
- list the factors that affect the rate of chemical reaction;
- explain the effects of changes in temperature, concentration, pressure and surface area on the rates of a chemical reaction;
- explain the effect of catalysts on the rate of chemical reaction;
- carry out an activity on how the factors affect the rate of chemical reaction;
- define the terms reversible reaction and irreversible reaction;
- define chemical equilibrium;
- describe the characteristics of chemical equilibrium;
- write the expression for the equilibrium constant for a reversible reaction;
- state Le Châtelier's principle;
- use Le Châtelier's principle to explain the effect of changes in temperature; pressure and concentration of reaction on equilibrium.

Forward Planning

Read the whole content on rates of reactions and chemical equilibrium (section 4.8). Prepare a plan of your own based on the topics, activities and experiments you will treat in each period so that you can complete the entire contents of the section within ten periods.

In your plan, indicate the duration of time allotted for the students to discuss activities, present their opinions, perform experiment, and also to harmonize concepts, gapped lectures and for other activities you perform in each period.

In this sub-unit, there are four experiments (Experiments 4.9, 4.10, 4.11 and 4.12). Prepare the chemicals, apparatuses and other materials required to perform the experiments. You better carry out the experiments yourself before you allow students to do them. Prepare other teaching aids like graphs and charts relevant for teaching this section.

Teaching Aids

- Charts and graphs.
- Refer to the student's text for the chemicals and apparatus required for performing Experiments 4.9, 4.10, 4.11 and 4.12.

Subject Matter Presentation

Reaction Rate

It is better to use group discussion, experiment and gapped lecture for this section.

It is advisable to begin the class with Activity 4.12. The activity helps students realize what reaction rate is. Let the students discuss it in groups and then ask them to present their findings to the class. Next, hold a whole-class discussion. After the discussion, harmonize it by presenting these ideas:

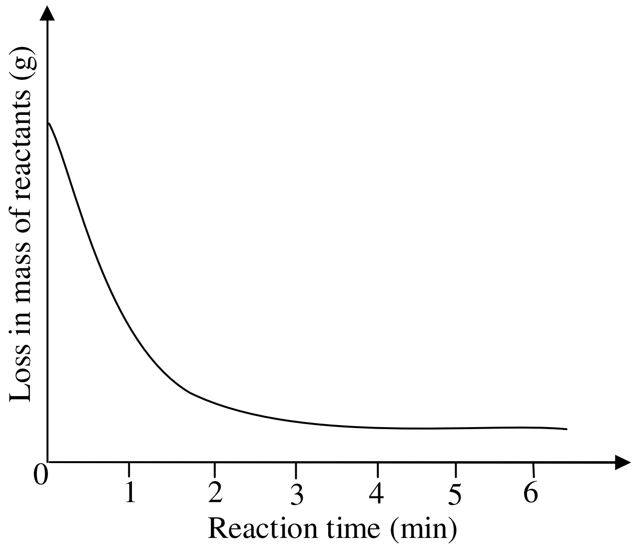
1. The rate of a reaction depends on the nature of reactants, concentration of reactants, temperature and catalysts. Assuming the other conditions are constant, some reactions are rapid at room temperature because such reactions are favored at this temperature and result in effective collisions between reactants. The combustion of gasoline is considered, fast reaction, and the rusting of iron is an example of slow reaction.

2. In hot tea. Generally, the higher the temperature, the faster the dissolution. This is because water molecules collide with the sugar more frequently at a higher temperature and as a result they dissolve the sugar easily.

Let students do experiment 4.9 in groups as given in the students textbook. Let them write down their observations and report to the class. Harmonize the concepts suggested by students with the following points:

Experiment 4.9

1. During the reaction, the mass gradually decreases due to evolution of CO_2 gas into the air.
2. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2 \uparrow$
3. Record the actual data from the experiment and plot a graph of loss in mass-versus-time, as illustrated in the following diagram.



4. At the start of the reaction, there is a high concentration of reactants (HCl and $CaCO_3$). Thus the reaction rate is maximum, and the graph becomes steep at this point. As the reaction proceeds, the amount of HCl and $CaCO_3$ available for the reaction decreases, and the rate of reaction slows down. The curve becomes less and less steep. Finally, the graph becomes a horizontal line at the end of the reaction. No more HCl and $CaCO_3$ are used up for the reaction at this point.

Introduce the important terms related to rate of chemical reaction and chemical equilibrium. Describe reaction rate, using the graphs you brought to the class. Explain the preconditions necessary for a chemical reaction to occur. Continue by presenting the factors that affect rate of chemical reaction.

Factors affecting the rates of chemical reactions

After completing the topic of the preconditions for a chemical reactions to occur, let the students discuss Activity 4.13 in groups for a few minutes and discover the factors influencing the rates of chemical reactions. Have some groups present their opinions to the class. Then harmonize their group discussion with the following facts:

1. The burning of charcoal is made faster by increasing the amount of air or charcoal, as rate depends on concentration of reactants.
2. The rate of combustion of a block of wood increases if we blow more air into the burning wood.

After harmonizing the concepts developed during Activity 4.13, let the students discuss Activity 4.14 in groups for a few minutes to discover how the nature of substances affects rate of reaction. Then ask some groups to present their findings to the class. Then, harmonize concepts:

Because reaction rate depends on the nature of reactants, the speed of combustion decreases in this order for these substances: kerosene, paper, then charcoal and last copper in air.

Continue your discussion by explaining how the factors like nature of reactants, temperature and concentration of reactants affect the rate of chemical reactions. Let the students discuss Activity 4.15 in groups for a few minutes. Then harmonize their group discussion with the following conclusions:

The rate of the reaction becomes slower with the additions of water because the rate of reaction is affected by concentration.

Let the students do experiments 4.10 in groups, as given in the student's textbook. Let the students write down the results of their observations and report to the class. Finally, harmonize concepts with the following points:

Experiment 4.10

1. In (b). If we compare the rate of evolution of the gas bubbles, we find that evolution of gas is faster with the more concentrated solution of H_2SO_4 (5M H_2SO_4).
2. $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2 \uparrow$
3. An increase in the concentration of the reactants provides more reacting species for collisions and results in an increase in the rate of the reaction.

Present the way that how the surface area of reactants affects the rate of reaction by letting them do Experiment 4.11. Have the students write reports about their observations and then ask some groups to present their findings to the class. Then, harmonize their opinions with the truth, as follows:

Experiment 4.11

1. Reaction (b) is faster.
2. When reaction occurs between the HCl and $CaCO_3$, many more contacts (collisions) take place between the powdered $CaCO_3$ and the HCl than between the lump of $CaCO_3$ and HCl. Therefore, we can conclude that an increase in surface area of the reactants increases the rate of the reaction.

Following Experiment 4.11, introduce students to the effects of catalysts on the rate of reactions. After that, let them discuss Activity 4.16 in groups for a few minutes and then have some groups present their opinions to the class. After their presentations, harmonize concepts:

Consider the melting of ice. In the presence of heat, ice melts and changes to liquid immediately. In this case, the conversion of ice (solid) to liquid is a physical change. But this physical change is facilitated by temperature.

Let the students do experiment 4.12 in groups, as given in the student's textbook. Let the students write down their observations and report it to the class. Make sure that their reports on the experiment coincide with the following points:

Experiment 4.12

1. In number (2). When MnO_2 is added, H_2O_2 decomposed rapidly into H_2O and O_2. This is because MnO_2 is acting as a catalyst and speeds up the decomposition of H_2O_2.
2. Oxygen gas.
3. Hydrogen peroxide decomposes very slowly on standing, but the process is made very much faster by adding a catalyst, MnO_2.

Chemical Equilibrium

Start teaching this topic by asking students to suggest their opinions about reversible and irreversible reactions. After getting feedback from them, inform them of the differences between the two. Also, tell them why we use single and double arrows when writing chemical equations for irreversible and reversible reactions, respectively. Be sure they are aware of the presence of opposing reactions in all reversible reactions. These are the forward and the reverse reactions. Explain that the rate of forward reaction is higher at the beginning and gradually decreases, while that of reverse reaction is smaller at the beginning and gradually increases with increasing reaction time. When the rates of the forward and reverse reactions are equal, the reaction is said to be in dynamic equilibrium.

Introduce them to the conditions required for a system to be in a state of dynamic equilibrium in terms of concentrations of reactants and products, and in terms or rates of forward and reverse reaction and progress of reaction.

Then, derive the mathematical expression for the law of chemical equilibrium, using examples. Let students practice writing equilibrium constant expressions for some reversible reactions. Finally show them how to calculate the equilibrium constant for a given reversible reaction and let them practice.

After introducing the chemical equilibrium of a reaction, continue by presenting Activity 4.17. This activity is designed to help students discover the importance of chemical equilibrium. So, let the students discuss Activity 4.17 in groups and then have them present their findings to the class. After the discussions, harmonize them by presenting these ideas:

The major practical importance of chemical equilibrium is in the chemical industry. In the chemical industry, high yield of product is obtained by adjusting external factors, such as temperature, concentration, and pressure.

Factors that Affect Chemical Equilibrium

Start this topic by letting students discuss Activity 4.18 for a few minutes in groups. Invite one or two groups to present their findings to the class. After the presentations, harmonize concepts.

Then, introduce the factors that affect chemical equilibrium and state LeChatelier's principle. Next, explain the effect of each factor on the chemical equilibrium. Tell them how factors like temperature, pressure and concentration affect the state of chemical equilibrium. To explain the effect of temperature, first explain how they can identify which reaction is exothermic and which endothermic, using some examples. Also emphasize that in a reversible reaction, if the forward reaction is endothermic, the reverse reaction will be exothermic and vice-versa. Let them also practice predicting the effect of temperature on the equilibrium of some reactions.

Next, continue introducing the effects of pressure or volume. Give some examples and let them also practice predicting the effect of pressure changes. Inform student that they should compare the number of moles (molecules) of gaseous substances on the reactant and product sides to predict the effect of pressure or volume change. Finally, tell them how concentration change can affect equilibrium. Also, explain that catalysts do not affect the state of chemical equilibrium, but only help reactions to reach this state at a faster rate.

Assessment

Assess each student's work throughout Section 4.8. To do this, you need to follow-up the participation of each student during the teaching learning process. Using student's performance list, you may record how each student:

- involves in discussing Activity 4.12 – 4.18.

- participates in presenting ideas of the group to the rest of the class.
- takes part in performing Experiments 4.9 – 4.12.
- involves in presenting to the rest of the class outcomes or observations from the experiments.
- answers questions raised during gapped lectures.
- performs activities given to them between the gapped lectures.

You may also give exercises 4.15 and 4.16 as homework and a test on the section. Correct their work and record their achievements. From all these records, make sure that most of the students have accomplished the competencies suggested for this section. Praise students working above the minimum requirement level and also give them extra activity. For students working below the minimum requirement level, it is preferable if you arrange extra lesson time to assist them catch up with the rest of the class.

Additional Questions

* 1. Suppose two reacting molecules collide with each other. Under what conditions do the colliding molecules not react?

2 Explain why the rates of most chemical reactions decrease over time.

* 3. Why does changing the volume (or pressure) of the reaction vessel have no effect on the following equilibrium?

$$CO(g) + Fe_3O_4(s) \rightleftharpoons CO_2(g) + 3FeO(s)$$

4. Define homogenous equilibrium and heterogeneous equilibrium. Give two examples for each.

5. 5 When equilibrium is shifted to the right, what happens to the concentration of reactants and products? Increase or decrease? Explain.

6 Name two properties that can be used to determine whether a chemical system is in equilibrium or not.

6. 7. How would each of the following changes affect the equilibrium position of the system?

$$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g) + 114.4KJ$$

a. adding HCl
b. cooling the system
c. adding a catalyst
d. removing Cl_2 from the system
e. decreasing the volume
f. removing all the reactants from the system
g. decreasing the pressure

Answers to Additional Questions

1. The colliding molecules could not react with each other if:
   - The colliding molecules are in poor orientation to each other.

- The activation energy of the reactants is insufficient.

2. Most reaction rates slow over time because, as reactants are consumed, fewer particles collide and therefore reaction slows.
3. In chemical equilibrium, there is no significant effect of pressure (volume) on the solid or liquid states of the substances.

   In the following equilibrium:

   $$\underbrace{CO(g)}_{\text{1 mole}} + Fe_3O_4(s) \rightleftharpoons \underbrace{CO_2(g)}_{\text{1 mole}} + 3FeO(s)$$

   only the number of moles CO on the reactant side and CO_2 on the product side affect pressure on the system. Therefore, there is no effect of pressure or volume on the system, since they have equal numbers of moles.
4. A chemical equilibrium in which all reactants and products are in the same phase is known as **homogeneous equilibrium**.

   Example : $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   If all the reactants and products are not in the same phase, the equilibrium is said to be **heterogeneous equilibrium**.

   Example : $2ZnS(s) + 3O_2(g) \rightleftharpoons 2ZnO(s) + 2SO_2(g)$
5. When equilibrium is shifted to the right, the concentration of the reactant decreases and the concentration or amount of product increases.
6. At the state of chemical equilibrium:
   - The rates of the forward and reverse reactions are equal.
   - There is no change in the concentration of reactants and products. (But this does not mean that the concentrations of the reactants and products are equal).
7. $$\underbrace{4HCl(g) + O_2(g)}_{\text{5 moles}} \rightleftharpoons \underbrace{2H_2O(g) + 2Cl_2(g)}_{\text{4 moles}} + 114.4\ KJ$$

   From this chemical equilibrium, it can be deduced that the reaction is exothermic and also all the substances are in the gaseous state (homogenous equilibrium).

   a. The system shifts to the right.
   b. The system shifts to the right.
   c. No effect on the position of equilibrium.
   d. The equilibrium shifts to the right.
   e. The equilibrium shifts to the right.
   f. The equilibrium shifts to the left.
   g. The equilibrium shifts to the left.

Answers to Exercises

Exercise 4.15

1. For a reaction to occur, the reacting molecules need sufficient activation energy, and also the collisions of molecules need to be in the proper orientation.
2. The rate of a reaction changes with time. This is because the concentration of reactants changes as the reaction proceeds.
3. a) $K_{eq} = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$

   b) $K_{eq} = \dfrac{[CH_3Cl]\,[HCl]}{[CH_4]\,[Cl_2]}$

   c) $K_{eq} = \dfrac{[SO_3]^2}{[SO_2]^2\,[O_2]}$

   d) $K_{eq} = \dfrac{[CO_2]}{[CO]}$

Exercise 4.16

For the reaction given in question 4.16, if the temperature is increased and the pressure is decreased, the equilibrium of the system shifts as follows:

a) If the temperature is increased

	Energy changes	The equilibrium shifts to the:
1	Exothermic	left direction
2	Endothermic	right direction
3	Endothermic	right direction
4	Exothermic	left direction
5	Endothermic	right direction
6	Exothermic	left direction

b) If the pressure is decreased

	Number of moles on the		The equilibrium shifts to the:
	Reactant side	Product side	
1	2	2	no effect of pressure
2	2	3	right direction
3	3	4	right direction
4	3	2	left direction
5	1	2	right direction
6	2	1	left direction

Answers to Review Exercise on Unit - 4

Part I

1. True	2. False	3. False	4. True	5. True
6. False	7. True	8. False	9. True	10. False

Part II

11. + 6
12. Temperature
13. Surface area of the reactants

Part III

14. 0.89 g
15. a) NH_3 is the limiting reactant.
    b) 44.1 g NO
16. 75.9%
17. The equilibrium of the system will shift to the:
    a) right
    b) right
    c) right
    d) left
    e) left
    f) left

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Unit 5 The Physical State of Matter

Unit Overview

Total Period Allotted 20

Unit 5 has five main parts. **Section 5.1** is the Introduction. It briefly describes the properties of gases, liquids and solids.

Section 5.2 introduces the kinetic molecular theory of matter and the properties of matter. The properties of gases are also presented in this section.

Section 5.3 deals with the gaseous state. It begins with the kinetic molecular theory of gases and the gas laws. Gas laws and related variables such as temperature, pressure, volume and number of moles are presented in detail.

Section 5.4 deals with the properties of liquids in terms of boiling, condensation, evaporation and vapor pressure. It also explains how the boiling point of a liquid is determined.

Section 5.5 deals with the solid state. Melting point, freezing point, sublimation, determination of melting point and phase changes are presented in this section.

The methodologies suggested for teaching in this unit are group discussion, gapped-lecture, demonstration, experiment, answer and question and drawing a map as methodologies.

Unit Outcomes

After completing this unit, students will be able to:

- understand the kinetic molecular theory and the properties of the three physical states of matter;
- know the behavior of gases using the variables volume, temperature, pressure and number of moles;
- know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, and molar heat of solidification;
- understand gas laws;

- develop skills in solving problems to which the gas laws apply;
- perform activities to illustrate gas laws;
- carry out experiments to determine the boiling points of liquids and the melting points of solids;
- perform an experiment to show phase changes;
- demonstrate scientific inquiry skills: observing, predicting, comparing and contrasting, measuring, interpreting data, drawing conclusions, applying concepts and making generalizations.

Main Contents

5.1 Introduction

Periods Allotted 1

Competencies

After completing this section, students will be able to:

- name the three physical states of matter.

Forward Planning

Read the contents in the introduction part from the students' text to recall the physical states of matter. Also read the teacher's guide to get more information about the start-up activity and Activity 5.1, and the methodology you implement to teach this section. You also need to plan how to budget your time for the activities you perform while teaching this section.

Subject Matter Presentation

It is better to use question and answer and discussion methods to teach this section.

You are advised to begin the class with the start up activity given in the student's textbook. The startup activity is designed to assist students know the existence of three states of matter and their phase changes. Help the students to form groups and let them discuss the activity. Next, invite a student to present the idea of his or her group to the class. Finally, organize a whole-class discussion. After the discussion, harmonize it by presenting these ideas:

Keep the tin cans containing ice as given in the student's text book. Let students observe what happens in each case. Help students to form groups and discuss their observations. The ice kept at room temperature remains solid for a long time, and after some time it melts slowly. The ice which is heated gently melts to form liquid water and further heating of the liquid water will change it to vapor state. Ask students how these phase changes occur. After harmonizing concepts that develop out of the startup activity, present Activity 5.1, which is intended to help them recall the three physical states of matter.

Have them discuss this activity. Guide and encourage them to take part in the group discussion and let them reflect the ideas of their discussion. Then, conclude the discussion by using the following example:

1. a) Stone, common salt, iron, and gold are examples of solid. We can see and touch them.
   b) Water, oil, coca cola are examples of liquid. We can see and touch them. They take up the shape of their container.
   c) Carbon dioxide, oxygen, argon are gases.
2. When an ice cube is heated, it melts to a liquid, and further heating evaporates it to a gas.

Here, apply brainstorming methodology to relate the topic with our daily lives. Let them discuss the following points again.

- make your students to list down some substance they know which is/are relevant to their life and exist as gas, liquid or solid.
- make them discuss the uses of the substances mentioned to our life.
- make also the following as a point of discussion. Is the absence of these substances has an impact on our life?
- then give them chance to present the result of their discussion and record what they suggest on the blackboard. Then, harmonize their discussion by giving the following examples.
  1. Oxygen is a gas. If we don't get oxygen we can die immediately.
  2. Water is a liquid. Lack of water has many impacts on our life. Generally life is impossible without water.
  3. Salt (NaCl) is a solid. It also gives a proper physiological functioning. If we do not take NaCl, the physiological functions of our body will face a problem.

Next, ask the students to define matter and to name the states of matter. Let the students write their answers in their individual notebooks. Check their work to assess their understanding. Then consolidate their knowledge by giving them appropriate answers to your questions.

Ask students to describe the properties of solid, liquid and gaseous states of matter. Let students write the answers in their note books. When they finish, check their work. Then continue your explanation of the properties of solid, liquid and gaseous states of matter. Finally introduce the plasma state of matter. Plasma has properties differing form those of the other states of matter. There are no complete atoms in plasma and no chemical changes can occur in it. It is of special interest to space scientists.

Assessment

You may assess each student's work by asking oral questions related to the physical states of matter. You can also give them Exercise 5.1 as class work or homework and check their work. See whether or not the suggested competencies are achieved by most of the students.

Answers to Exercise 5.1

1. Oxygen can exist as a liquid and a solid at very low temperatures and high pressures.
2. Let students hold discussions in groups about the three states of matter in terms of space between their particles, distance between them, volume and shape, and compressibility and density.
3. Dry ice is solid carbon dioxide.

5.2 Kinetic Theory and Properties of Matter

Period Allotted 3

Competencies

***After completing this section, students will be able to*:**

- give examples for each of the three physical states of matter;
- state the kinetic theory of matter;
- explain the properties of matter in terms of the kinetic theory;
- compare and contrast the three physical states of matter.

Forward Planning

Read the contents in the students' text in section 5.2 to get information about the kinetic theory of matter and about properties of gases, liquids and solids. Prepare a plan of your own that shows which topics and activities you are going to deal with during each period in order to complete the whole section within three periods. In your plan, include the time allotted for group discussion, presentation, gapped lecture and other activities you may perform in each period. Also read the teacher's guide to have knowledge the methodologies you implement to teach this section and to get more information about Activities 5.2, 5.3 and 5.4.

Subject Matter Presentation

It is better to use discussion and gapped lecture to teach this section.

After introducing the topic of the section, let the students discuss Activity 5.2 in groups. The activity is designed to help students realize the movements of particles, strength of the forces and distance between particles in the three state of matter.

Guide them in their approaches to this activity. Let them present the ideas of their discussion to the class. After the discussion, harmonize it by presenting these ideas:

1.

Particles	Motion of molecules	Distances between molecules	Attraction between molecule
Sodium (solid)	low	very close to each other	high
Bromine (liquid)	medium	close to each other	medium
Chlorine (gas)	high	very far apart	low

2. In the gaseous state, the molecules possess the highest kinetic energy, since the motion of particles of gases is higher as compared to those of liquids and solids.

Ask students if they have any ideas about the kinetic theory of matter. Encourage them to respond. Following their responses, continue to introduce the assumptions that underlie the kinetic theory of matter.

Properties of matter

After you complete the kinetic molecular theory of matter, start with Activity 5.3 to proceed to the next part. The activity is designed to assist students discover how the density of the same substance can vary with its stat. Let the students discuss Activity 5.3 in groups for a few minutes. Ask some of the groups to present concepts related to the points in the activities. Then, harmonize their ideas with the following.

1. The density of a substance in the gaseous state is lower, compared to the density of the same substance in the liquid or solid state. That is because the particles are much farther apart in the gaseous state. At normal atmospheric pressure, most liquid substances are denser than gases. The higher density of solids results from

the fact that the particles of a solid are more closely packed together than those of a liquid or a gas.

2. During compression, gas particles that are initially very far apart are brought close together. Hence, the volume of a given sample of a gas can be highly decreased. Liquids are much less compressible than gases because liquid particles are more closely packed together. Solids can be considered incompressible. After harmonizing the concepts that the students developed during Activity 5.3, introduce them the properties of gases, liquids and solids in detail. Next, let them discuss Activity 5.4 for a few minutes in groups, discover the effect of increasing amount of gas on its volume and then ask one or two students from different groups to suggest their opinions to the class. After that, harmonize the concepts.

When air is blown into a balloon, the volume of the balloon increases because the air in the balloon takes up space.

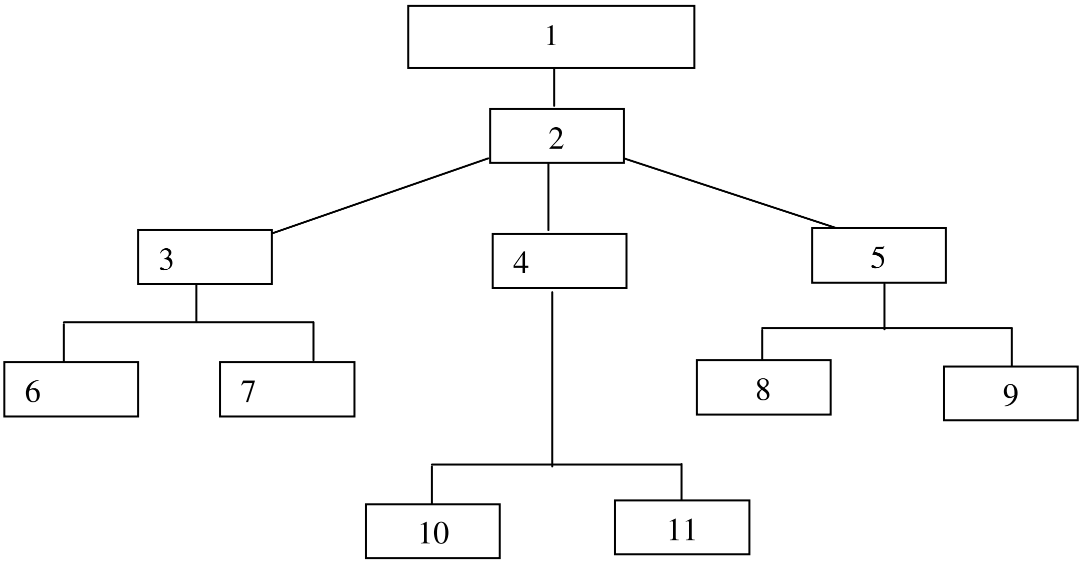
Finally, explain the properties of gases, liquids and solids. During this presentation, ask the students to suggest the kinetic theory explanation of why the states of matter exhibit particular properties.

Concept Mapping

After you introduce the physical states of matter, list the following terms on the blackboard and let some volunteer students try to construct the concept map using the terms:

- state
- physical properties
- incompressible
- solid
- liquid
- tightly packed particles
- gas
- virtually incompressible
- compressible
- particles far apart
- closely packed particles

You can harmonize, the concept suggested by the student using the following map.



1. physical properties
2. state
3. solid
4. liquid
5. gas
6. incompressible
7. tightly packed particles
8. compressible
9. particles far apart
10. virtually incompressible
11. closely packed particles

Assessment

Assess each student's work throughout section 5.2. To do this, follow-up the performance of each student and record what you observe related to how the student:

- involves in group discussions.
- takes part in presentations.
- performs activities given between gapped lecture.
- answers questions during harmonizing concepts or stabilization.

You can give them exercise 5.2 as class work or homework and check their work. Based on the achievements of each student, see whether or not the suggested competencies for the section are achieved. Appreciate students working above the minimum requirement level. Give the necessary assistance to those working below the minimum requirement level.

Additional Questions

1. How does the kinetic energy of particles vary as a function of temperature?
2. Use the kinetic-molecular theory to explain why both gases and liquids are fluids.
3. Containers A, B and C contain sample of H_2O at 130°C, 20°C, and –15°C, respectively, all at atmospheric pressure.
   a. Which sample has molecules with the greatest kinetic energy?
   b. Which sample has the greatest density?
   c. Which sample has a definite volume?
   d. Which sample has a definite shape?
   e. Which sample is the most compressible?
   f. Which sample has the most regular arrangement of molecules?

Answers to Additional Questions

1. Kinetic energy and temperature of a gas are directly related. Temperature is a measure of the average kinetic energy of the particle in a sample of matter. At a given temperature, all gases have the same average kinetic energy.
2. This is because both liquids and gases can easily flow.
3. a. C   d. A

   b. A   e. C

   c. A and B   f. A

Answers to Exercise 5.2

A. Gas < Liquid < Solid.

B. Gas < Liquid < Solid.

C. Solid < Liquid < Gas.

D. Solid < Liquid < Gas.

5.3 The Gaseous State

Period Allotted 11

Competencies

After completing this section, students will be able to:

- explain the assumptions of the kinetic molecular theory of gases;
- describe the properties of gases, using the kinetic molecular theory;
- describe the behavior of gases by using the variables V (volume), T (Temperature), P (pressure) and n (number of moles);

- state Boyle's law;
- perform an activity that shows changes in volume and pressure of gases to illustrate Boyle's law;
- apply Boyle's law in solving problems;
- state Charles' law;
- perform an activity that shows changes in volume and temperature of gases to illustrate Charles' law;
- apply Charles' law in solving problems;
- derive the combined gas law equations from Boyle's law and Charles' law;
- use the combined gas law to calculate changes in volume, pressure or temperature;
- define an ideal gas;
- derive the ideal gas equation from Boyle's law, Charles' law and Avogadro's law;
- compare the nature of real gases with ideal gases;
- solve problems based on ideal gas equations;
- define diffusion;
- state Graham's law of diffusion;
- carry out an activity to compare the rate of diffusion of two different gases;
- apply Graham's law of diffusion in solving problems.

Forward Planning

Read thoroughly from the student's text the contents on the gaseous state of matter. This will help you to understand the details on the kinetic molecular theory of gases and the gas laws and to make the appropriate plan. Design a plan of your own that shows the topics, activities and experiments you are going to deal with during each period in such a way that the entire contents in the section can be covered within eleven periods. In your plan, indicate the duration of time allotted for the students to discuss an activity, to make presentations, harmonize concepts, for gapped lectures and other activities you perform in each period.

There are three experiments in section 5.3. These are Experiments 5.1, 5.2 and 5.3. Prepare the materials, chemicals and apparatuses required to perform these experiments. Carry out the experiments yourself before you allow the students to perform them. Plan how to manage and assist students when they perform the experiments. Read the teacher's guide on this section to get more concepts about Activity 5.5 – 5.14 and also to get information about which methodologies you implement to teach this section.

Teaching Aids

Refer to the materials required to perform Experiment 5.1, 5.2 and 5.3 from the students text.

Subject Matter Presentation

You better use group discussion, demonstration, experiment and gapped lecture for this section.

First summarize the main points of the previous lesson and introduce them to the topic of the following lesson to be the kinetic molecular theory of gases.

First introduce the topic of the section. Then, continue with the suggested activity which is designed to help students discover the effect of temperature on the volume of gases. Have students discuss activity 5.5 a few minutes and present their conclusions to the class. Guide students in writing short paragraphs and let them present the ideas of their discussion to the class. You may get different answers from different students. Finally, harmonize their ideas as follows:

More air is recommended. This is because, at cold temperatures, the air is highly compressed, and more air is needed to fill the tire.

After that, continue to introduce the basic assumptions of the kinetic molecular theory of gases. Then, ask the students to review the assumptions.

Next, continue teaching the gas laws. Start the lesson with Activity 5.6. The activity is designed to assist students discover the effect of temperature on pressure of gases. First, let the students discuss activity 5.6 in groups. Next, let them present the ideas of their discussion. Then conclude their discussions.

a) The pressure in the car tire increases on a hot day because the pressure of a gas is directly proportional to the temperature.

b) This is due to the expansion of the gas inside the bulb.

After harmonizing concepts in Activity 5.6, present the definitions and units of the variables used to describe the behaviour of gases (V, T and P) and also define molar volume. Next, proceed by teaching Boyle's law.

Boyle's Law

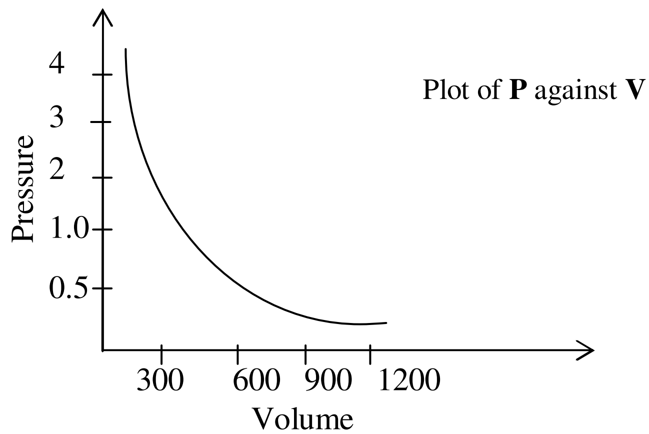
Start the lesson using an activity. Activity 5.7 is designed to help students discover the effect of pressure on the volume of gases. Let the students discuss Activity 5.7 in groups for a few minutes. Let them present the ideas of their group discussions to the class. Finally, harmonize their ideas as follows.

The helium weather balloon expands as it rises more into the air because the air becomes less dense. As we move up, the atmospheric pressure decreases. As pressure decreases, the volume increases, and the balloon expands as it rises up.

Now, explain Boyle's law. State the law and introduce its mathematical expression. After you do so, solve some problems as examples. Give the students class work and homework and let them practice solving problems using this law.

Demonstrate experiment 5.1 to the class, as given in the student's textbook. Let the students record what they observe during the demonstrations and report their findings to the class.

Experiment 5.1



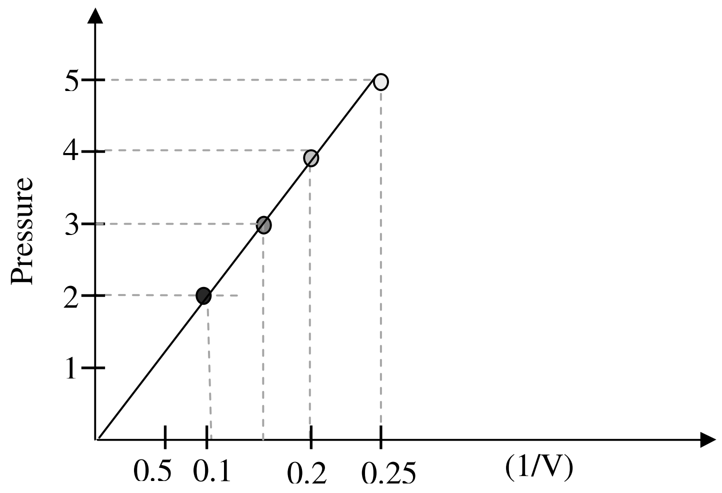
The graph shows a plot of P versus V for a sample of gas. The pressure increases with decreasing volume at constant temperature.

$$V \alpha \frac{1}{P}.$$

Let students discover the relationship between pressure and volume of a gas practically by drawing a graph.

Help them to form groups and let them do activity 5.8 for a few minutes and discuss what the graph looks like. Then conclude their discussions.

Pressure (atm)	1	2	3	4	5
Volume (mL)	20	10	6.7	5	4
$\frac{1}{V}$	0.05	0.1	0.149	0.2	0.25



This graph shows plot of P versus $\frac{1}{V}$ for the same mass of gas at constant temperature. The straight line indicates that volume varies inversely with pressure.

Charles' Law

Start the lesson on this topic with Activity 5.9. The activity is designed to help students realize the relationship between temperature and volume of gas.

Let students discuss activity 5.9 for a few minutes and then present their conclusions to the class. Help the students to write short paragraphs and let them present the ideas of their discussion to the class. You may get different answers from different students. Finally, harmonize their ideas as follows:

1. Popcorn is a type of corn that explodes when the kernel is heated. Corn is able to pop because, unlike other grains, its kernel has a hard moisture-sealed hull and a dense starchy filling. This allows pressure to build inside the kernels until an explosive "pop" result.
2. As temperature of a gas increases, its volume correspondingly increases at a constant pressure. This property can be explained by the kinetic-molecular theory. At a higher temperature, gas particles move faster, striking each other and the walls of their container more frequently and with greater forces.

Now, explain Charles' law and introduce its mathematical expression. Solve some problems as examples. Give exercises as class work and homework and check their work. Make sure that the students understand how to apply Charles' law to solve problems.

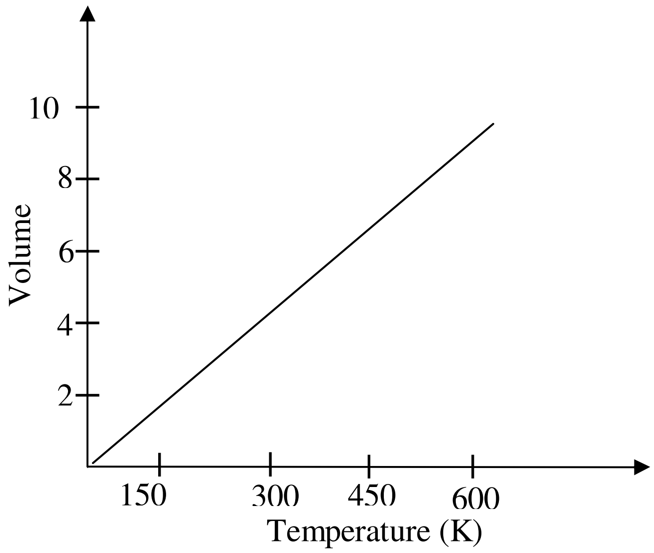
Let students perform Experiment 5.2 in groups as given in the student's textbook. Let the students write down what they observe during the experiment and report to the class. Check whether or not their reports coincide with the following points.

Experiment 5.2

1. As the flask is heated, the air molecules inside the flask expand and move through the capillary tube and form bubbles in the beaker of water.
2. The experiment is an illustration of the increase in the volume of a gas with an increase in temperature.

After the experiment continue with the suggested activity. This activity is designed to assist students discover the relationship between temperature and volume by drawing graph. Let them discuss Activity 5.10 for a few minutes. Next, let them present the ideas of their discussion. Then, harmonize their ideas with the actual facts.

Volume of Nitrogen gas (L)	**Temperature (k)**
4.28	303
5.79	410
7.77	550



b) $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ Charles' law, at constant pressure

$V_1 = 4.28$ L $\quad T_1 = 303$ K

$V_2 = ?$ $\quad T_2 = 700$ K

$$V_2 = \frac{4.28\ \text{L} \times 700\ \text{K}}{303\ \text{K}}$$

$V_2 = 9.89$ L

c) As temperature increases, so does volume if pressure is kept constant. That is, the volume of a gas is directly proportional to the absolute temperature at constant pressure.

Then, introduce the mathematical expression of the combined gas law. Show them how to relate the pressure, volume and temperature of the same quantity of a gas under two different conditions. Use the combined gas law to solve some problems as examples. Let the students practice solving problems by using the combined gas law. Give them class work and homework, and check their work.

After you conclude the presentation of the combined gas law, present Avogadro's law**.** State the law and introduce its mathematical expression and the variables that should remain constant for this law to be valid.

You can continue with Activity 5.11. This activity enables students to discover the relationship between volume of gas and its amount. Then, let the student discuss Activity 5.11 for a few minutes. Let them present the ideas of their discussion to the class. After the discussion, harmonize it by presenting these ideas:

i. As you pump the ball, more and more air is added to it. As a result of this the ball becomes strong enough.

ii. The number of gas particles inside the ball increases.

iii. Avogadro's Law.

Ideal gas equation

Start the lesson with Activity 5.12. It is designed to help students to understand relationships between the variables P, V, n and T. First, let the students discuss Activity 5.12 in their groups for a few minutes. Then, have them present the ideas of their discussion to the class. Then conclude their discussion.

When a lot of air is added to a balloon, the number of molecules and the pressure inside the balloon increase. But the volume remains the same. Finally, the balloon bursts because its volume cannot increase any further with the increase in the number of molecules and pressure.

Now present the ideal gas equation. Show its derivation, explain the equation and the variables it relates. Solve some problems as examples. Finally, give some exercises as class work and homework. Check their work to see if the students can apply this law properly.

Graham's law of diffusion

Start the lesson with an activity. The activity is designed to help students identify the relationship between lighter and heavier gases with their rates of diffusion. Let them discuss Activity 5.13 for a few minutes. Next, let them present the ideas of their discussions to the class. Then, conclude their discussions.

A light gas diffuses faster than a heavy gas. Thus, the balloon filled with helium deflates, through its invisible holes, faster over time than does the balloon filled with air. This is because helium is lighter than air.

At the end of this section, introduce your students to Graham's law of diffusion. First, let them define what diffusion is and then explain the meaning of the term. Then, state Graham's law of diffusion and present the mathematical expressions to show the relationship between the rates of diffusion of two gases and their densities and molecular masses. Solve one or two problems as examples. Ask some questions and let the students practice solving problems related to Graham's law. Check and correct their work.

Let students perform Experiment 5.3 in groups, as given in the student's textbook. Let the students record what they observe during the experiment and report their findings to the class.

Experiment 5.3

This is a demonstration of diffusion of gases.

1. HCl gas moved the shorter distance to the white ring. This white ring in the tube is due to the formation of ammonium chloride when NH_3 reacts with HCl.
2. Generally, the rate of diffusion of two gases can be compared based on their molecular masses. Since the molecular mass of NH_3 is 17 g/mole, and that of HCl is 36.5 g/mole, NH_3 is lighter than HCl. As a result, NH_3 moves faster in the tube and forms a white ring closer to the HCl end of the tube.
3. NH_3 diffuses faster than HCl.
4. The lighter the gas, the faster the diffusion.

Next, have them discuss Activity 5.14 for a few minutes in groups, discover the effect of applying force or pressure on volume of a gas and let them present the ideas of the groups to a class. Then harmonize their discussion. If you apply pressure to the walls of a flexible container, gases can be compressed as a result of a decrease in volume. This occurs due to the larger space between the particles of gases. Gases also have a tendency to fill the whole container when heated. This shows the expansion of gases.

Assessment

Assess each student's work throughout Section 5.3. You can do this by watching carefully how every student is doing the task given to him/her throughout the teaching-learning process on the gaseous state. Record the performance of each student in your student's performance list based on what you observe during your contact with students in each period. You can make records considering how every student:

- takes part in discussing Activities 5.5 – 5.14.
- involves in presenting ideas or concepts after discussion.
- participates in performing Experiment 5.1 – 5.3.
- takes part in presenting outcomes and observations about the experiments.
- answers questions raised during discussions, harmonizing concepts, gapped lectures and stabilization.
- performs the given activities between gapped lectures.

You can give them Exercises 5.3, 5.4, 5.5, 5.6, 5.7 as class work or homework as well as a test related to contents in section 5.3. Correct their work and record their achievements in the student's performance list. Based on the cumulative records you have, check whether or not the competencies suggested for section 5.3 are achieved by most of the students. Praise students working above the minimum requirement level and encourage them to continue working hard. Give them additional activity. In relation to students working below the minimum requirement level, help them by arranging extra lesson time or giving them additional exercise. You can give the questions suggested to them from the additional questions given in the teacher's guide. In addition, set questions on Boyle's, Charle's, the combined gas law, on the ideal gas equation and Graham's law if necessary and let them attempt the questions.

Additional Questions

* 1. A sample of gas occupies 0.5 L at 850 torr. What volume does it occupy at standard atmospheric pressure? (Temperature is constant)

2. By what factor must the pressure be increased if 10 liters of oxygen gas is compressed to 4 liters, assuming temperature remains constant?
3. At what temperature will the volume of a gas be doubled if its original temperature is 27°C, assuming pressure remains constant?
4. A sample of gas is heated from 27°C to 627°C. By what factor does the volume increase if the pressure remains constant?

* 5. A sample of gas has a pressure of 760 mmHg at 30°C. Calculate the temperature of the gas if the pressure is changed to 1.5 atm.

6. A gas exerts a pressure of 3 atm at 0°C. What will the pressure of the gas be if the original temperature is doubled, at a constant volume?

* 7. A gas collected at 127°C has a volume of 10 L and a pressure of 800 torr. What would the temperature of the gas be if the volume is changed to 7200 cm^3 at standard pressure?

* 8. A certain sample of gas occupies 25.5 L at 298 K and 153.3 kPa. Calculate its volume at STP.

* 9. Determine the volume of a 1.5g of hydrogen gas at STP.

10. A 20g sample of gas has a pressure of 4 atm and a volume of 1500ml at 30 °C. What is the molecular mass of the gas?

* 11. Arrange the following gases in the order of increasing rates of diffusion:

SO_3, CH_4, H_2, N_2O, Cl_2, O_3, He

12. What is the relative rate of diffusion of H_2 and N_2?

13. An unknown gas diffuses 1.27 times faster than Cl_2 under the same conditions.
    d. Is the unknown gas heavier or lighter than Cl_2?
    e. What is the molecular mass of the unknown gas?

Answers to Additional Questions

1. 0.56 L
2. By a factor of 2.5.
3. 600 K
4. By a factor of 3. ($V_2 = 3V_1$)
5. 454.5k
6. 6 atm
7. 273.6 K
8. 35.3 L
9. 16.8 L
10. 83.3 g/mole
11. SO_3, Cl_2, O_3, N_2O, CH_4, He, H_2
12. $r_{H2} = 3.74\, r_{N2}$
13. a. Lighter than Cl_2
    b. 44 g/mole

Answers to Exercises

Exercise 5.3

a) 760 mmHg = 760 torr = 76 cm Hg = 1 atm

$\therefore$ 500 mmHg = 50 cm Hg = 500 torr = 0.657 atm

b) 1L = 1000 mL = 1000 cm^3 = 10^{-6} m^3

1 mL = 10^{-3} L = 10^{-3} m^3

$\therefore$ 100 dm^3 = 100, 000 cm^3 = 100,000 mL = 100 L = 10^{-1} m^3

c) K = °C + 273

= 54 + 273 = 327 K

$^oF = \left(\frac{9}{5} \times {}^oC\right) + 32$

= 1.8 × 54°F + 32 = 129.2°F

Exercise 5.4

1. Initial conditions

$V_1 = 10\ m^3$

$P_1 = 100$ kPa

Final conditions

$V_2 = 20\ m^3$

P_2 = ?

Apply Boyle's law ($P_1V_1 = P_2V_2$) (*at constant temperature*)

$$P_2 = \frac{P_1 \times V_1}{V_2} = \frac{100\ kPa \times 10\ m^3}{20\ m^3} = 50\ kPa$$

2. V_2 = 14 L (*use the same steps as above*)

Exercise 5.5

1. Initial conditions

T_1 = -173°C = 100 K

$V_1 = V_1$

Final conditions

T_2 = 27°C = 300 K

$V_2 = V_2$

Apply Charles' law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$V_2 = \frac{T_2 \times V_1}{T_1} = \frac{300\,K}{100\,K} \times V_1 = 3V_1$$

2. a) Initial conditions

T_1 = 17°C = 290 K

$V_1 = V_1$

Final conditions

T_2 = ?

$V_2 = ½\ V_1$

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

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

$$T_2 = \frac{1/2V_1}{V_1} \times T_1 = \frac{1}{2} \times 290\,K = 145\,K = -128\,^0C$$

b) $V_2 = 2V_1$

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

$$T_2 = \frac{2V_1}{V_1} \times T_1 = 2 \times 290\,K = 580\,K = 307\,^0C$$

c) $V_2 = 3V_1$

$$T_2 = \frac{3V_1}{V_1} \times T_1$$

$$T_2 = 3T_1 = 3 \times 290\,K = 870\,K = 597\,^0C$$

3. $T_1 = 25^oC = 298$ K $T_2 = 100^oC = 373$ k

$P_1 = 1$ atm $P_2 = 1$ atm

$V_1 = 1.5\ dm^3$ $V_2 = ?$

Apply the combined gas law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

$$V_2 = \frac{1\,atm \times 1.5\,dm^3 \times 373\,K}{1\,atm \times 298\,K} = 1.88\ dm^3$$

Exercise 5.6

$$V_2 = \frac{60\ \text{kPa} \times 50\ \text{cm}^3}{308\ \text{K}} \times \frac{273\ \text{K}}{1.013 \times 10^2\ \text{kPa}} = 26.24\,\text{cm}^3$$

Exercise 5.7

d = 1.77 g/L

P = 1.34 atm

T = 0.0821 L atm K^{-1} mol^{-1}

$$PV = nRT \text{ and } n = \frac{m}{M} \Rightarrow pv = \frac{mRT}{M}$$

$$\text{and } M = \frac{dRT}{p}$$

$$\therefore \quad M = \frac{1.77\ g/\cancel{L} \times 0.0821\cancel{L}.\ \cancel{atm}.\ \cancel{K}^{-1}\ mol^{-1} \times 303\ \cancel{K}}{1.34\ \cancel{atm}}$$

$$= 32.\ 86\ g/mole$$

5.4 The Liquid State

Period Allotted 3

Competencies

After completing this section, students will be able to:

- explain the terms: evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization and molar heat of condensation;
- carry out an activity to demonstrate the concept of vapor pressure;
- carry out an activity to determine the boiling points of water and ethanol.

Forward Planning

Read the contents from the student's text about the liquid state thoroughly. This will help you to make a plan how you can present the contents. You need to design a plan that shows the contents, activity and experiment you are going to deal with during each period so that you can cover the entire contents in the section within three periods. In your plan, indicate the time allotted for every activity you will perform in each period. Also read the teacher's guide on this section to get more information about the suggested activities, observations and analysis of the experiments and about the methodologies you implement.

There are two experiments in this section (Experiment 5.4 and 5.5). Prepare the materials required to perform the experiment and also plan how to manage students when the students perform the experiments and discuss activities.

Teaching Aids

Refer to the students' text for the apparatuses and chemicals required to perform Experiments 5.4 and 5.5.

Subject-Matter Presentation

It is advisable if you use group discussion, experiment and gapped lecture as your methodologies for this section.

After introducing the topic, start the lesson with Activity 5.15. The activity is designed to help students realize the main cause why some liquids are volatile and others do not and also discover why liquids boil at low temperature at higher altitudes. Help students to form groups. Next, let them discuss Activity 5.15 in groups for a few minutes and then present the ideas of their discussion. Then harmonize concepts.

1. This is because some liquids evaporate readily at room temperature, whereas others do not. This difference is explained in terms of forces of attraction between their particles. Volatile liquids have relatively weak forces of attraction between their particles, whereas nonvolatile liquids have stronger forces between their particles.
2. As altitude increases, atmospheric pressure decreases. Therefore, if you boil water at higher altitudes, it boils at lower temperatures because the external air pressure decreases as altitude increases.

Next, introduce the intermolecular forces in liquids and then continue by teaching energy changes in liquids. First, let the students discuss activity 5.16 in groups and present their findings to the class. Next, hold a whole-class discussion. Then conclude the discussion.

As the steam from the hot water comes in contact with the cold mirror, it gets cold and changes to liquid water.

In presenting this section, basic terminology and concepts related to liquids should be emphasized. This includes evaporation, condensation, and vapor pressure.

Help students form groups and have them discuss each of the following basic concepts: Boiling, boiling point, the dependence of boiling point on altitude, normal boiling point, heat of vaporization, condensation, molar heat of condensation and evaporation.

Next let them discuss Activity 5.17 in groups for a few minutes and then have one or two groups present their views. Then, harmonize concepts.

a. As you move up (higher altitude) air becomes less dense. Therefore, a gas at a lower altitude with higher density exerted strong pressure. On the other hand, at a higher altitude gases becomes less dense and exert less pressure.

b. The temperature at which a crystalline solid is converted to a liquid is known as the melting point. The melting point of a substance can fluctuate by the presence of impurities. The melting point of an ice occurs below 0°C in the presence of impurities.

Evaluate your students as to how they differentiate between boiling and evaporation. Next, introduce the concept of vapor pressure. Explain that a liquid boils when its saturated vapor pressure is equal to the atmospheric pressure.

Your presentation should include the determination of vapor pressure and the boiling points of water. Let the students perform **Experiment 5.4** in groups and present what they observe in the experiment to the class. Harmonize concepts:

Experiment 5.4

As the water in the flask is heated, its molecules escape to create pressure above the surface of the liquid. This is called vapor pressure of the liquid. The vapor pressure of the liquid pushes the liquid to a certain height in the U- tube**.**

Following Experiment 5.4, let the students perform **Experiment 5.5** in groups and present what they observe in the experiment to the class. Harmonize concepts:

Experiment 5.5

The temperature rises steadily as the water is heated. The boiling point of water actually depends on altitude and atmospheric pressure.

- Let the students record the thermometer reading.
- Once boiling starts, the temperature of the water remains constant (at its boiling point).
- Porcelain chips are added as boiling chips. The purpose is to avoid the sudden bumping of the boiling liquid.

Assessment

You can assess each student's work throughout Section 5.4. You can do this through strict follow-up of each student's performance and by recording in the student's performance list. You can make a record based on how each student:

- involves in discussing Activities 5.15, 5.16 and 5.17.
- participates in presenting views and concepts after discussions.
- takes part in performing Experiments 5.4 and 5.5.
- involves in presenting outcomes, observations and analysis of the experiments.
- does activities given to him or her during gapped lectures
- answers questions raised during harmonizing concepts and stabilization.

From your record, make sure that the suggested competencies for Section 5.4 are achieved by most of the students. Appreciate students working above the minimum requirement level. Assist those working below the minimum requirement level by giving them additional exercise. Set questions related to the concepts on the liquid state.

5.5 The Solid State

Period Allotted 2

Competencies

After completing this section, students will be able to:

- explain the terms: melting, fusion, sublimation, melting point, freezing point, molar heat of fusion and molar heat of solidification;
- describe phase changes;
- explain temperature changes associated with phase changes;
- determine the melting point of ice;
- perform an experiment to show the phase changes from ice to liquid water and then to water vapor.

Forward Planning

Read the contents from the student's text on the solid state of matter. Plan which contents and activities you will treat during each period so that you can cover all the contents in the section within two periods. In addition, indicate the time allotted to every activity you will perform in each period.

Subject Matter Presentation

It is advisable to use group discussion and gapped lecture as methodologies to teach this section.

After introducing the topic of the section, let the students discuss Activity 5.18 in groups for a few minutes. Let them present the ideas of their discussion to the class. After the discussion, harmonize it by presenting these ideas:

1. When a crystal of iodine is warmed, it vaporizes directly to the gaseous state. The direct change of a substance from the solid to the gaseous state is known as sublimation. Therefore, iodine undergoes sublimation.
2. When ethyl alcohol is put in an open container, it disappears after some time because it is a volatile substance.

To apply brainstorming methodology give them the following activity.

What is the significance of cooling liquid substance for our life. Make them the rational of cooling liquid!

- List down their response/answer on the blackboard.
- Confirm the correct ideas they emanate.

Compare their ideas with the following:

Food kept under cool condition stay for a long time than food kept under hot condition. Under a cooled condition food do not produce micro organism which may decompose food. The food can stay fresh for a long time.

Help students form groups and have them discuss concepts and terms related to the topic after you list them on the blackboard. Following the discussion, let students reflect on their ideas and summarize them. They should be able to account for the properties of solids, using the kinetic theory of matter and understand phenomena such as melting, melting point, freezing, freezing point, molar heat of fusion and sublimation, and energy changes.

In your school laboratory, demonstrate phase change and the determination of melting point of solids, using melting ice. This has to be supported by the heating curve given in the student text. Explain to the students why the temperature remains constant at the melting point of a solid and at the boiling point of a liquid. Inform your students that, at these temperatures (at the melting point of a solid and the boiling point of a liquid), the temperature remains constant because the heat energy is used up to break the intermolecular forces in the solid and liquid states.

Assessment

Assess each student's work throughout section 5.5. To do this, see how every student participates in discussion, takes part in presentation after discussion, attempts the activities given to him/her between the gapped lectures, answers questions raised during the gapped lectures or during harmonizing concepts. You can also ask oral questions related to melting point, heat of fusion, solidification and sublimation. From their responses and from what you observe, see whether or not the suggested competencies for section 5.5 are achieved. Praise students working above the minimum requirement level and give necessary assistance to those working below the minimum requirement level.

Answers to Review exercise on unit 5

Part I		Part II Matching	
1. solid	6. liquid	11. D	15. I
2. gas	7. gas	12. H	16. E
3. solid	8. gas	13. K	17. A
4. gas	9. solid	14. G	18. C
5. gas and liquid	10. gas		

Part III Fill in the Blank

19. Vibrate
20. Gas, liquid
21. Freezing point
22. External pressure
23. Melting point
24. - 273.15^{o}C
25. Avogadro's

Part IV Short Answers

26 The difference between volatile and nonvolatile substances or liquids is presented as follows:

Volatile liquids	Nonvolatile liquids
1. They evaporate readily at room temperature.	1. They have a low tendency to evaporate at room temperature.
2. They have weak intermolecular forces between their particles.	2. They have relatively strong intermolecular forces between their particles.
3. They have high vapor pressure. Examples: ethanol, benzene, diethyl ether, etc.	3. They have low vapor pressure. Examples: sulphuric acid, water, molten ionic compounds.

27. During the melting process, heat is absorbed by the system, and the temperature remains constant at the melting point. The heat added is used to overcome the attractive forces between the molecules in the solid.

28. The rate of evaporation of a liquid depends on the following three factors:

 1. temperature
 2. intermolecular forces
 3. surface area of the liquid

Part V Problems to Solve

29. a) 0.95 atm
 b) 950 mmHg
 c) 542 torr
 d) 98.9 Kpa
 e) 7 atm

30. a) 69.6 K
 b) 32.2 atm

31. 44.8 L
32. 0.57 L
33. 0.90 g
34. 112 ^{o}C
35. 216 s
36. 1.9 atm
37. a) 0.69 L b) 61 atm

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Federal Democratic Republic of Ethiopia
Ministry of Education

Chemistry Syllabus and Minimum Learning Competencies, Grade 9

2009

MINIMUM LEARNING COMPETENCIES (MLCs)

No	Areas of Competencies	Minimum Learning Competencies (MLCs)
1	SUBSTANCES	• Describe Dalton's and Modern Atomic theory • Compare and contrast Dalton's and Modern Atomic theory • Write the charges and the masses of the three fundamental subatomic particles • Explain the terms- atomic number, mass number, atomic mass and isotopes • Determine the number of protons, number of electrons, and number of neutrons from atomic number and mass number • Name the five atomic models and state Bohr's postulates • Explain Energy levels, Valence electrons, and electron configuration • Write the ground state electron configuration for given elements and represent them diagrammatically • Define chemical bonding and explain why atoms form chemical bonds with other atoms • Define ionic bond, describe its formation and explain the general properties of ionic compounds • Define covalent bond, describe its formation and explain the general properties of covalent compounds • Make models of molecules to show single, double and triple bonds using balls and sticks of locally available materials • Give examples of simple ionic and covalent compounds and draw their electron dot structures • Explain polarity of covalent molecules and distinguish between polar and non polar molecules • Explain coordinate covalent (dative) bond formation using examples • Define metallic bonding and describe thermal and electrical conductivity of metal in relation to metallic bond • Define intermolecular forces, explain dipole-dipole and dispersion forces and illustrate using examples • Explain the effects of hydrogen bond and dispersion forces on the properties of substances • Name and give examples for the three physical states of matter • State kinetic theory of matter , explain and compare the properties of the three physical states of matter in terms of kinetic theory • State and explain Boyle's law, Charles' law, combined gas law, Avogadro's law and Graham's law of diffusion and do calculations to which the laws apply • Perform activities to show the changes in temperature, pressure and volume of gases to illustrate Boyle's and Charles' laws • Explain the terms evaporation, condensation, vapor pressure, boiling point, heat of vaporization and heat of condensation • Explain the terms- melting, fusion, sublimation, melting point,

		freezing point, heat of fusion and heat of solidification • Carry out activities to demonstrate the concepts of vapor pressure and to determine the boiling points of water and ethanol • Describe phase changes and explain temperature changes associated to phase changes • Demonstrate an experiment to show phase changes using ice, liquid water and water vapor • Demonstrate Scientific inquiry skills:- observing, predicting, classifying, comparing and contrasting, making model, communicating, measuring, asking questions, interpreting data, drawing conclusions, applying concepts, relating causes and effects and making generalizations
2	**CHEMICAL REACTIONS**	• Define Chemical reaction and give examples • State the laws of definite proportion and the law of conservation of mass and illustrate with examples • Demonstrate the law of conservation of mass using simple experiment • Balance chemical equations using the inspection and the Least Common Multiple (LCM) methods • Discuss energy changes in chemical reactions, distinguish between exothermic and endothermic reaction and illustrate using diagram • Appreciate the importance of chemical changes in the production of new substances and energy' • Identify four types of chemical reactions and give examples • Deduce mole ratios from balanced chemical equations • Solve mass-mass problems based on the given chemical equations • State Avogadro's principle, solve Volume-Volume and mass-volume problems based on balanced chemical equations • Determine excess and deficient (limiting substances) in a reaction. • Describe the percentage, actual and theoretical yields of a chemical reaction • Calculate the percentage yield of a reaction from the given information • Define redox reaction, the terms oxidation and reduction in terms of electron transfer and give examples • Define oxidation number and determine the oxidation number of an element in a given formula(in a molecular, an ion or a compound) • Describe oxidizing and reducing agents and distinguish between them • Analyze a given redox reaction by specifying the substance reduced, the substance oxidized, the oxidizing agent and the reducing agent • Distinguish between redox and non-redox reaction

		• Explain the reaction rate and describe an activity to illustrate it • Describe how collision, activation energy and proper orientation a chemical reaction to occur • List and explain the factors that affect reaction rate and describe briefly an activity to illustrate each • Define reversible reaction, irreversible reactions and chemical equilibrium • Describe the characteristics of chemical equilibrium • Write an expression of equilibrium constant of a reversible reaction • State the Lechatlier's principle and explain factors affecting chemical equilibrium • Demonstrate scientific enquiry skills:- observing, inferring, predicting, classifying, comparing and contrasting, communicating, measuring, asking • questions, designing skills, interpreting data, drawing conclusions, applying concepts, relating causes and effects and problem solving
3	**CLASSIFICATION IN CHEMISTRY**	• Describe Periodicity • State the modern periodic law • Define the terms Period and Group • Explain the relationship between the electronic configuration structure of the modern periodic table • Identify and explain the three classes and the four blocks of the elements of the periodic table • Tell the number of the main groups and sub groups and give specific names for the main group elements • Tell the number of periods and classify them as short, long and incomplete based on the number of elements they contain • Predict from the given atomic number of an element its position in the periodic table • Deduce the properties of an element from its position in the periodic table. • Explain the general trends in properties of elements down a group and across a period of the periodic table • Make a model to demonstrate the trends in properties of elements in the periodic table • Appreciate the importance of the classification in the study of chemistry • Demonstrate the scientific enquiry skills:- observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusions, applying concepts and problem solving.

Introduction

Students need an understanding of chemistry to fully appreciate the complexity and interactions present in their world. Chemistry deals with composition, properties and transformation of substances. The study of chemistry should help the students to increase their analytical and perceptive capabilities and provide opportunity to explore vocations in the field of chemistry. It also provides students with the knowledge and skills in chemistry and technology and enables them to solve problems and make decisions in everyday life based on scientific attitudes and noble values.

To achieve the above mentioned purpose and to cop up with the current demands of chemical knowledge for technological, agricultural and industrial development chemistry curriculum has to be updated all through the grade levels.

The chemistry curriculum for grades 9 and 10 are then revised based on the new curriculum framework of Ethiopian schools taking into consideration the assessment made in March 2008 in selected secondary schools of the country, international experiences and the current situations of our country.

Feedback of the assessment made it clear that the secondary school curriculum also has a problem of content overload, content difficulty inappropriate to the grade levels, giving less emphasis to active learning methods, limitations in integrating agriculture and technology and unnecessary repetitions. The curriculum revision then addressed these problems basically.

The revision also considered international standards. In this regard chemistry curriculum of different countries including India, Malaysia, Singapore, England, Ghana, and Uganda were taken as references. The international consultant has also contributed in keeping standards by sharing experiences and involving in the revision processes.

Consideration of the above mentioned major points led to some restructuring of the units of grade 9 and 10. In grade 9, the titles Chemical Symbols, Formulas and Equations and chemical Reaction are collectively given under the unit Chemical Reaction and stoichiometry. Similarly, Hydrocarbons and Oxygen Derivative Hydrocarbons are given in grade 10 as Introduction to Organic Chemistry and some parts of the oxygen derivative hydrocarbons are taken to grade 12. In the unit of Electrolysis of grade 10 Electrolysis of aqueous solutions and their quantitative aspects are also taken to grade 12. To solve the problems of content overload, chemistry of only three selected metals (Al, Fe, and Cu) is treated in grade 10. Besides, some agricultural and industrial applications are integrated in the units.

The format of the syllabus is different from the traditionally used format. There are only three columns of competency, content and 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. The statement **"minimum requirement level"** should not mislead and should be understood as the **"standard level"**. Students working at the standard level are expected to achieve the competencies set for the grade level successfully. Teachers should give special considerations for those who are working above and below the standard levels by encouraging the ones that work above the standard and by giving extra attention for those who work below the standard. Enrichment activities should be designed and provided through optional exercises, supplementary exercises, etc, for those who work above and remedial activities should be designed and provided through tutorial, peer tutorial, group work, etc, for those who work below.

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 and what is to be assessed. The environment and culture of the students must also be considered.

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 techniques suit the particular purpose of assessment.

Correlating Instruction, Evaluation, and 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 & Tests
Indirect	• Concept Mapping/Formation/ Attainment • Inquiry • Problem Solving	• Individual/Group: Presentations • Oral Assessments • Performance Assessments • Written Assignments
Experiential	• Conducting Experiments • Field Observations & 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

For the better implementation of the curriculum material the chemistry syllabus of each grade level is accompanied by materials such as flowchart, minimum learning competencies, student textbook, teachers guide, practical activities manual and student workbook.

This syllabus for grades 9 and 10 were revised by 13 teachers and national and international education experts.

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Ato Wasihun Bitew (Somali)

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Allotment of Periods
for Units and Sub-units of Chemistry
Grades 9

Grade	*Unit*	*Sub-unit*	*Number of Periods*	
			Sub-unit	*Total*
9	**Unit 1: Structure of the Atom**	1.1 Atomic theory 1.2 Discoveries of the fundamental subatomic particles and the atomic nucleus 1.3 Composition of an atom and the isotopes 1.4 The atomic models	3 3 3 6	15
	Unit 2: Periodic Classification of the Elements	2.1 Introduction 2.2 The modern periodic table 2.3 Periodic properties in the periodic table 2.4 Advantages of periodic classification	1 5 6 1	13
	Unit 3: Chemical Bonding and Intermolecular Forces	3.1 Chemical bonding 3.2 Ionic bonding 3.3 Covalent bonding 3.4 Metallic bonding 3.5 Intermolecular forces	1 3 8 1 4	17
	Unit 4: Chemical Reaction and Stoichemistry	4.1 Introduction 4.2 Fundamental laws of chemical reactions 4.3 Chemical equations 4.4 Energy changes in chemical reactions 4.5 Types of chemical reactions 4.6 Stoichemistry 4.7 Oxidation and reduction reactions 4.8 Rate of chemical reaction and chemical equilibrium	1 2 3 3 3 10 5 10	32
	Unit 5: The Physical States of Substances	5.1 Introduction 5.2 Kinetic theory and properties of matter 5.3 The gaseous state 5.4 The liquid state 5.5 The solid state	1 3 11 3 2	20

General Objectives of Grade 9 Chemistry

To Develop Understanding and Acquire knowledge of:

- Dalton's atomic theory and modern atomic theory,
- Discoveries of the sub-atomic particles,
- The relationship between the sub-atomic particles and the atomic numbers, mass numbers, atomic masses and isotopes,
- Development of the atomic models and arrangement of the electrons in the atoms,
- Periodic classification of the elements and its importance in studying chemistry,
- The major types of chemical bonding and intermolecular forces,
- Types of chemical reactions and representing chemical reactions by chemical equations and
- The three physical states of matter.

To Develop Skills and Abilities of:

- Writing electron configuration for the elements,
- Drawing structures and make models for some molecules,
- Writing balanced equations and solving related problems,
- Design and conduct simple experiments relevant to their level and
- Applying gas laws' equations to solve gas variables (pressure, temperature and volume).

To Develop the Habit and Attitude of:

- Realizing that properties of substances are attributed to their atomic structures,
- Appreciating problem solving,
- Appreciating that all substances are the results of chemical combination of the atoms and
- Develop personality characteristics such as neatness, exactness, diligence, responsibility and carefulness.

Unit 1: Structure of the atom (*15 periods*)

Unit outcomes: Students will be able to:

- comprehend Dalton's atomic theory and modern atomic theory;
- understand the discovery of the electron and the nucleus;
- know the terms like atomic number, mass number, atomic mass, isotope, energy level, valence electrons and electron configuration;
- understand the Dalton, the Thomson, the Rutherford, the Bohr and the quantum mechanical atomic models;
- develop skills in
  - determining the number of protons, electrons and neutrons of atoms from atomic numbers and mass numbers,
  - calculating the atomic masses of elements that have isotopes,
  - writing the ground state electron configurations of atoms using sub-energy levels and drawing diagrammatic representations of atoms.
- demonstrate scientific inquiry skills: observing, comparing and contrasting, communicating, asking questions, and applying concepts.

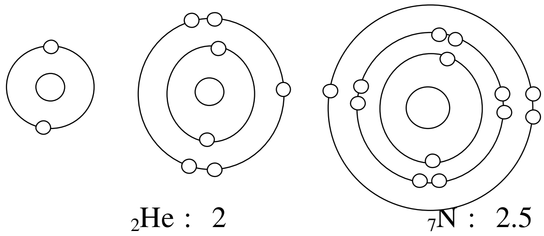
Competencies	***Contents***	***Suggested Activities***
Students will be able to: • Describe Dalton's atomic theory	**1. Structure of the atom** **1.1 Atomic theory** *(3 periods)* • Dalton's atomic theory	Students should appreciate that the idea of atoms as the building blocks from which all matter is formed was first suggested by the ancient Greeks although they had no evidence to support this theory. Students should know that in 1808 the scientist John Dalton proposed an atomic theory in which he suggested that: 1. Elements are made of small particles called atoms 1. Atoms can neither be created nor destroyed 2. All atoms of the same element are identical and have the same mass and size 3. Atoms of different elements have different masses and size 4. Atoms combine in small whole numbers to form compounds Students should discuss each point in Dalton's atomic theory and determine whether it is still valid today.
• Describe the modern atomic theory	• Modern atomic theory	Students should appreciate that at the time that Dalton's theory was proposed nothing was known about the internal structure of the atom. As a result of our increasing knowledge about atomic structure we now know that statement #2 and statement #3 are no longer true.

Competencies	Contents	Suggested Activities
• Compare and contrast Dalton's atomic theory and the modern atomic theory		Students should attempt to modify Dalton's statements in the light of modern knowledge about the atom. The modified statements could be: 1. Elements are made of small particles called atoms 2. Atoms cannot be created or destroyed during ordinary chemical reactions 3. All atoms of the same element have the same atomic number but may vary in mass number due to the presence of different isotopes 4. Atoms of different elements are different 5. Atoms combine in small whole numbers to form compounds Students could be shown the symbols that Dalton used to represent atoms of different elements and discuss why these symbols are no longer used.
	1.2 Discoveries of fundamental subatomic particles and the atomic nucleus *(3 periods)*	Students should appreciate that our modern understanding of atomic structure came about as a result of the discovery of its component particles.
• Explain the discovery of electron	• Discovery of the electron	Students should be aware that the electron was discovered by the physicist J.J. Thomson. Students should be able to give a brief summary of the famous experiment carried out by Thomson. This could include the following points: • The apparatus consisted of two electrodes in a tube from which most of the air had been removed • When a high potential difference was applied across the electrodes a stream of particles passed between them • Thomson called these particles cathode rays because they came from the cathode • The cathode rays moved straight to the anode, and through a hole in it, to a screen coated with zinc sulphide • The zinc sulphide glowed when struck by the cathode rays

Competencies	***Contents***	***Suggested Activities***
• Explain the terms atomic mass and isotope • Calculate the atomic masses of elements that have isotopes.	• Atomic mass, and isotope	Students should understand that in the atoms of an element: • the number of protons is always the same • the number of neutrons varies slightly Students should know that atoms of the same element which have different numbers of neutrons will have different mass numbers and are called isotopes. Students should appreciate that a natural sample of an element is likely to contain a mixture of two or more isotopes. In determining the atomic mass of the element. We must take into account that it is a mixture of isotopes with different mass numbers. Students should be aware that: • Some elements are composed almost entirely of one isotope therefore the atomic mass of the element will be the same as the mass number of the isotope – for example carbon exists as a mixture of the isotopes carbon-12, carbon-13 and carbon-14, but carbon-12 accounts for 98.9% of the element. • Some elements contain significant proportions of different isotopes and an accurate value for the atomic mass is obtained by taking the mass number and percentage of each isotope into account – for example, boron exists as a mixture of the isotopes boron-10, 20%, and boron-11, 80%. Therefore both isotopes make a significant contribution to the atomic mass of boron. The atomic mass is equal to the weighted average mass of its isotopes. Students should calculate the atomic mass of boron using the following method: $10 \times \frac{20}{100} + 11 \times \frac{80}{100} = 2 + 8.8 = 10.8$ Students could apply this method to find accurate values for the atomic masses of other elements such as copper and chlorine.

Competencies	Contents	Suggested Activities
• Name the five atomic models	**1.4 The atomic models** *(6 periods)*	Students should appreciate from the work already carried out in this unit that scientists picture of the nature of an atom has changed over the years as a result of new discoveries being made. Students should understand the important features of a number of atomic models proposed over the years.
• Describe the Dalton, • Thomson and Rutherford Model	• The Dalton atomic model • The Thomson atomic model • The Rutherford atomic model	Dalton model of the atom: • A small hard ball which is indivisible and cannot be destroyed Thomson model of the atom: • A positively charged solidly sphere • Electrons stuck uniformly in it like plums in a pudding Rutherford model of the atom: • A nucleus consisting of positively charged protons • Electrons moving around the nucleus
• State Bohr's Postulates • Describe the Bohr's model • Describe the quantum	• The Bohr atomic model - Bohr's postulates • The quantum mechanical	Bohr model of the atom: • A nucleus consisting of positively charged protons and neutrons which carry no charge • Electrons moving around the nucleus in a set of orbits • For any atom there is a fixed set of orbits possible • The energy of an electron remains the same as long as it stays in the same orbit • When electrons move between orbits they release or absorb energy • When electrons fall from a higher (excited) state to a lower (ground) state they give out fixed amounts of energy. But, when the electrons jump from a lower to a higher energy state they absorb fixed amounts of energy. Students should appreciate how the picture of the atom became more sophisticated as scientists gained knowledge about it. Students should appreciate that our modern

Competencies	***Contents***	***Suggested Activities***
mechanical model • Describe main energy level and sub energy level	 - Main energy level - Sub-energy level	picture of the structure of the atom is based on a branch of science called quantum mechanics. Students should know that quantum mechanics came about because scientists found that the tiny particles in an atom did not obey the classical laws of physics postulated by Sir Isaac Newton. Students should appreciate that in the quantum mechanics atomic model: • Electrons are located in orbitals • An orbital is a volume of space inside which there is a high probability of finding a specified electrons • Orbitals are represented as clouds of charge • Energy levels in atoms are arranged in series in which adjacent levels get closer moving up the series • With the exception of the first energy level, each main energy level has sub-levels associated with it Students should know that: • Energy levels are numbers 1, 2, 3, 4, etc. or letters K, L, M, N etc. • Energy sub levels are given the letters s, p, d, f • Each type of sub level has a set number of orbitals: - s – 1 orbital - p – 3 orbitals - d – 5 orbitals - f – 7 orbitals • Each orbital can have a maximum of 2 electrons • The positions of the electrons around the nucleus of an atom in orbitals and sub orbitals is described as the electronic configuration of the atom
• Define the term electronic configuration • Write the	- Electronic configuration	Students should know the order in which orbitals are filled and that this corresponds to placing electrons in the lowest energy levels available: 1s, 2s, 2p, 3s, 3p, 4s, 3d. etc Students should note that the 4s orbital is filled before the 3d orbital because it is

Competencies	***Contents***	***Suggested Activities***
• ground state electronic configuration of the elements • Draw diagrams to show the electronic configuration of the first 18 elements		lower in energy. Students should be able to write the ground state electron configurations of the elements a suitable form such as: • Hydrogen – $1s^1$ • Helium – $1s^2$ • Lithium – $1s^2$, $2s^1$ etc. Students should be able to draw diagrams to show the electron configurations of the first 20 elements. Students should consider some examples. This could include the diagrammatic representation of the electronic configurations of the elements He, N & Al  $_2$He : 2 $_7$N : 2.5 $_{13}$Al : 2, 8, 3 Students could draw the electronic configurations of other elements.
• Write the electronic configuration of the elements using sub energy levels • Write electronic configuration of the elements using noble gas as a core		Students should write electronic configuration of the elements using sub energy levels. They should also use short cut for electron configuration of elements with large atomic numbers using noble gas as a core.
• Describe valence electrons	- Valence electrons	Students should know that the electrons in the outermost orbital have the highest energy and are called valence electrons. Students should appreciate that valence electrons are: • furthest from the nucleus of the atom • the most easily lost responsible for the chemistry of the element

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: comprehend Dalton's atomic theory and modern atomic theory; explain the discovery of electron, neutron and nucleus; write the relative charges, relative masses and absolute masses of electron, proton and neutron; determine the No of neutrons of an element from given values of atomic number and mass number; explain the terms atomic mass and isotopes; calculate the atomic mass of elements that have isotopes, name the five atomic models and state Bohrs postulates, describe main energy level and sub energy level; write electronic configuration of elements and show them diagrammatically and describe valence electrons.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 2: Periodic classification of elements *(13 periods)*

Unit outcomes: Students will be able to:

- Understand the periodic classification of the elements.
- Develop the skills of correlating the electron configuration of elements with the periodicity of the elements, predicting the trends of periodic properties of elements in the periodic table.
- Appreciate the importance of classification in chemistry.
- Demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusion, applying concepts and problem solving.

Competencies	***Contents***	***Suggested Activities***
Students will be able to: • Describe periodicity	**2. Periodic classification of the elements** **2.1 Introduction** *(1 period)*	Students should understand that in classifying elements, scientists were guided by the similarities in chemical properties. The elements are arranged in a table in such a way that elements with similar chemical properties appear at regular intervals or periods. Students could revise early attempts to classify elements such as Dobereiner's Triads and Newland's Octaves.
• State Mendeleev's Periodic law • State modern periodic law	**2.2 The modern Periodic Table** *(5 periods)* • The Periodic law	Students should be aware that the modern Periodic Table is based on the work of the Russian chemist Dmitri Mendeleev. Students could research the work of Mendeleev, his foresight in leaving a gap between silicon and tin, and his predictions about the properties of the missing elements eg. Germanium.
• Describe period • Describe group	• Groups and periods	Students should know that the rows of the Periodic Table are called periods. They should understand that elements in the same period: • Have the same number of main energy levels • Increase in atomic number by one unit passing across the period • Decrease in metal properties passing from left to right Students should know that the columns of the Periodic Table are called groups. They should understand that elements in the same group: • Have the same number of electrons in the outermost shell • Have similar chemical properties
• Explain the relationship between the electronic	• Classification of the elements	Students should compare the electronic configurations of the first 20 elements with their positions in the Periodic Table. They should see that: • Elements whose valence electrons are in s orbital

Competencies	***Contents***	***Suggested Activities***
configuration and the structure of the modern Periodic Table		appear to the left-hand side (in Groups 1 and 2) • Elements whose valence electrons are in p orbital appear to the right-hand side (in Groups 3 to 8)
• Describe the three classes of the elements in the modern Periodic Table	- The representative elements (s-and p-blocks)	Students should understand that this can be used to classify elements in blocks with the Periodic Table e.g. • s-block elements are those in Groups 1 and 2 • p-block elements are those in Groups 3 to 8
• Explain the four blocks of the elements related with their electronic configuration in the modern Periodic Table • Tell the block of an element from its electronic configuration	- The transition metals (d- block)	Students could write the electronic configurations for elements with atomic numbers 21 to 36 of the Periodic Table. From this they will see that another block of elements emerges in which the valence electrons are in d orbital. Students should appreciate that d-block elements appear between the s-block and p-block elements. Students should appreciate that: • the d-block elements are sometimes called the transition metals • the d-block contains four series of elements corresponding to valence electrons in the 3d, 4d, 5d and 6d
	- The rare earths (f-block)	Students should be aware that there are two series of f-block elements corresponding to valence electrons in the 4f and 5f orbital. These series are collectively called the rare earths
• Give group names for the main group elements		Students should be aware that some groups have traditional names and others as family of the first member in the group. which are often used in textbooks: Group 1 – alkali metals Group 2 – alkaline earth metals Group 3 - boron family Group 4 - carbon family Group 5 - nitrogen family Group 6 - oxygen family Group 7 – halogens Group 8 – noble gases Students should be aware that some periods are long while others are short and are in incomplete. Students should be able to give examples of each

Competencies	***Contents***	***Suggested Activities***
• Classify the periods into short, long and incomplete periods • Tell the number of groups and periods in the modern Periodic Table • Tell the number of elements in each period • Predict the period and group of an element from its atomic number • Tell the block and group of an element from its electronic configuration		Students should recall the number of elements each period of the Periodic Table. Students should be able to: • Identify an element from its group and period • Identify the group and period of an element from its atomic number • Identify the block to which an element belongs from its electronic configuration. Students should be given information about elements and from the information, identify the element and its position in the Periodic Table.
• Explain the general trends in properties of elements down a group of the Periodic Table	**2.3 Periodic properties in the Periodic Table** *(6 periods)* • Periodic properties within a group - Nuclear charge - Atomic radius - Ionization potential - Electron	Students should appreciate that all of the elements in a group have the same number of electrons in the outer shell, so the elements will have some similarities, but different numbers of main energy levels so they will also have some differences. Students should consider the following characteristics of the elements in Group 1. They should discuss and explain any trends which are apparent. • Nuclear charge – number of protons in the nucleus • Atomic radius – half the distance between the nuclei of adjacent atoms in a substance • Ionization potential – the energy needed to form an ion from an atom by removing valence electron(s) (to form a positive ion) • Electron affinity – the ease with which an atom will accept an electron (to form a negative ion) • Electro negativity- the ability of an atom, when it

Competencies	***Contents***	***Suggested Activities***
	affinity - Electro negativity - Metallic character	is in a molecule, to attract or pull electrons towards itself • Metallic character – the tendency to lose electrons. A similar study should be carried out on the elements in other groups. Students should compare the trends in Group 1 or 2 (a group of metals) with those in Group 7 (a group of non-metals). Students could carry out practical work that exemplifies trend in reactivity down a group e.g. • Reaction of Group 2 metals with water • Displacement reactions involving Group 7 halogens and metal halide solutions Students should be able to use their knowledge of the trends in reactivity passing down a group to make predictions about the reactivity of elements. Students should understand whether these properties increase or decrease down a group and able to give reasons.
• Explain the general trends in properties of elements across a period of the Periodic Table	• Periodic properties within a period - Nuclear charge - Atomic radius - Ionization potential - Electron affinity - Electro negativity - Metallic character	Students should appreciate that all of the elements in a period will have the same main energy level but an increasing number of electrons in the outer shell. Students should consider the following characteristics of the elements in Period 2. They should discuss and explain any trends which are apparent. • Nuclear charge • Atomic radius • Ionization potential • Electron affinity • Electro negativity • Metallic character A similar study could be carried out on the elements in Period 3. Students should compare the trends across each period.
• Deduce the properties of an element from its position in the Periodic Table		Students should use their knowledge about group and periodic trends to make predictions about the physical and chemical properties of some unfamiliar elements. They could then research the properties of these elements and evaluate how accurate their predictions were.

Competencies	***Contents***	***Suggested Activities***
• Make a chart to show the trends in properties of elements in the Periodic Table		Students should make some form of chart to show the trend in properties of a group or period of elements.
• Describe the advantages of the periodic classification in the study of chemistry	**2.4 Advantages of periodic classification** *(1 period)*	Students should be able to use their knowledge of classification of elements to - read names, symbols, atomic numbers and atomic masses - state period and group of element - deduce number of shells, valence electrons, behaviour of an element and the nature of the oxides of element. Students should appreciate that the Periodic Table puts elements with similar properties into groups with gradual change in reactivity. In periods, some changes are gradual while others are abrupt. Students should be aware that the Periodic Table allows us to make predictions about the properties of elements. This can also be extended to their compounds. Students could be given data on some compounds of elements in the same group and asked to identify any similarities in physical and chemical properties. They could make predictions about similar compounds of the same elements. Students could consider and attempt to explain the diagonal relationship, based on similar chemical properties, between lithium and magnesium.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: Describe periodicity; state Mendeleev's periodic law and the modern periodic law; describe period and group; explain the relationship between electronic configuration of the elements and the structure of the periodic table; describe the three classes and the four blocks of elements in the periodic table; give group names for the main group elements; classify periods into short, long and incomplete periods; tell the number of groups and periods in the periodic table and the number of elements in each period; predict the period and group of an element from its atomic number; explain the general trends in properties of elements down a group and across a period of the periodic table; deduce the properties of an element from its position in the periodic table and describe the advantages of the periodic classification of the elements.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

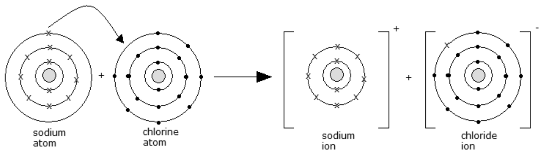
Unit 3: Chemical bonding and intermolecular forces *(17 periods)*

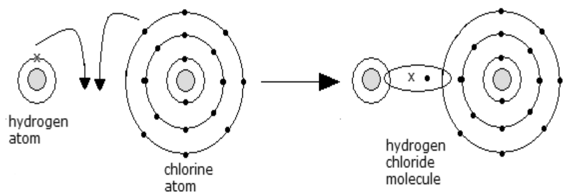
Unit outcomes: Students will be able to:

- discuss the formation of ionic, covalent and metallic bonds;
- know the general properties of substances containing ionic, covalent and metallic bonds;
- develop the skills of drawing the electron dot or Lewis structures for simple ionic and covalent compounds;
- understand the origin of polarity within molecules;
- understand the formation and nature of intermolecular forces;
- appreciate the importance of intermolecular forces in plant and animal life;
- demonstrate scientific inquiry skills: observing, predicting, making model, communicating, asking questions, measuring, applying concepts, comparing and contrasting, relating cause and effects.

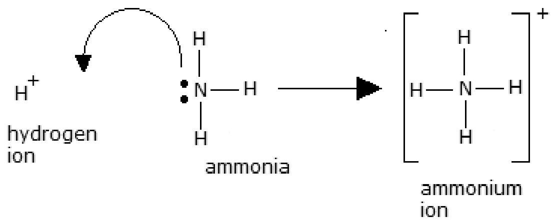
Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**3. Chemical bonding and inter- molecular forces**	
• Define chemical bonding • Explain why atoms form chemical bonds	**3.1 Chemical bonding** *(1 period)*	Students should appreciate that a chemical bond is any force of attraction between two particles. The particles may be atoms, ions or molecules. Students should understand that atoms form bonds in order to achieve a more stable electron arrangement – a full outer shell of electrons. This can be achieved by: • Transferring one or more electrons – ionic bonding • Sharing one or more pairs of electrons – covalent bonding Energy changes occur during bond formation – these usually result in the final substance being more stable. Students should be aware that atoms which already have a full outer shell of electrons – the elements in Group 8 (noble gases) – have little chemistry as they do not readily form bonds.
• Explain the term ion • Illustrate the formation of ions by giving examples	**3.2 Ionic bonding** *(3 periods)* • Formation of ionic bonding	Students should know that an ion is an atom which has lost or gained one or more electrons. • Metals tend to lose electrons to form positively charged ions or cations • Non-metals tend to gain electrons to form negatively charged ions or anions

Competencies	***Contents***	***Suggested Activities***
• Define ionic bonding • Describe the formation of an ionic bond • Give examples of simple ionic compounds • Draw Lewis structures or electron-dot	• Lewis Formulas of ionic compounds	• Hydrogen forms both a cation, H^+ (hydrogen ion), and an anion, H^- (hydride ion) Students should be able to relate the position of elements in the Periodic Table to the normal ion formed by them: • Group 1 – M^+ e.g. Na^+ • Group 2 – M^{2+} e.g. Mg^{2+} • Group 3 – M^{3+} e.g. Al^{3+} • Group 5 – X^{3-} e.g. N^{3-} • Group 6 – X^{2-} e.g. O^{2-} • Group 7 – X^- e.g. Cl^- Students could be asked to write the formulas of the ions formed by atoms of different elements. Students should appreciate that the easiest way for a metal to achieve a full outer shell of electrons is by losing electrons e.g. • sodium atom (2.8.1) → sodium ion (2.8) The alternative to this would be for the sodium atom to gain 7 electrons which would be very unstable. Similarly the easiest way for a non-metal to achieve a full outer shell of electrons is by gaining electrons e.g. • chlorine atom (2.8.7) → chloride ion (2.8.8) The alternative to this would be for the chlorine atom to lose 7 electrons which would require a large amount of energy and produce an unstable ion. Students should be aware that ionic bonding is the result of forces of attraction between oppositely charged ions. Students should be able to show how an ionic bond is formed by the transfer of one or more electrons from a metal atom to a non-metal atom. Students should represent the formation of an ionic compound using Lewis structure. An example of dot and cross diagram (Lewis formula) for sodium and chlorine to form

Competencies	***Contents***	***Suggested Activities***
formulas of simple ionic compounds		sodium chloride could be given as follows.  (2,8,1) (2,8,7) (2,8) (2,8,8) Students should also be able to represent the Lewis structure using symbols. eg. $Na + Cl \rightarrow Na^{+} [Cl]^{-}$ Students should appreciate that, when forming ions, the name of a non-metal takes the ending –ide e.g. nitride, oxide, chloride. Students could name the ionic compounds formed from given metals and non-metals and draw Lewis structures to show the ionic bonding.
• Explain the general properties of ionic compounds.	• General properties of ionic compounds	Students should discuss general properties of ionic compounds including: • crystalline nature • high melting points and boiling points • ability to conduct an electric current when molten or in aqueous solution • solubility in polar solvents like water Students should be aware that solubility in water in itself is not a proof that a compound is ionic. There are ionic compounds which are effectively insoluble in water, and covalent compounds like glucose which are very soluble.
• Investigate the properties of given samples of ionic compounds.		Students could be given samples of ionic compounds and asked to investigate their properties.
	3.3 Covalent bonding *(8 periods)*	
• Define covalent bonding • Describe the formation of a	• Formation of Covalent bond	Students should know that covalent bonds are formed when atoms share pairs of electrons. Students should understand that the bond arises due to the electrostatic attraction between the negative electrons and the

Competencies	*Contents*	*Suggested Activities*
covalent bond • Draw Lewis structures or electron-dot formulas of simple covalent molecules • Give examples of different types of covalent molecules	• Lewis formula of covalent molecules.	positively charged nuclei of the two atoms. Students should appreciate that covalent bonds are generally formed between atoms of non-metals. Students should be able to show how a covalent bond is formed by the sharing of a pair of electrons. One electron is donated by each non-metal atom. The shared pair of electrons is considered to exist in the outer orbital of both of the atoms. Students should represent the formation of a covalent molecule Lewis structure Symbols or diagrams can be used to illustrate it. Eg. H^x + Cl  Students should appreciate that the gaseous elements in Groups 5, 6 and 7 do not exist as atoms but as molecules in which the atoms are joined by covalent bonds e.g. N_2, O_2, Cl_2. Students should practice drawing Lewis structures for common covalent compounds. These could include hydrogen chloride, water, ammonia and methane. Students should appreciate that in some covalent compounds, atoms are joined by double and triple covalent bonds: • C-C e.g. ethane • C=C e.g. ethene • C≡C e.g. ethyne There are also double and triple bonds between atoms in some molecules of gases e.g. • N≡N in nitrogen • O=O in oxygen Students should practice drawing Lewis structures for common covalent compounds in which there are double and triple bonds. Students could research other covalent compounds in which there are double and

Competencies	*Contents*	*Suggested Activities*
• Make models of covalent molecules to show single, double and triple bonds using sticks and balls or locally available materials. • Explain polarity in covalent molecules • Distinguish between polar and non polar covalent molecules	• Polarity in covalent molecules.	triple bonds e.g. C=S, C≡N. Students should make models to show covalent compounds in which there are single bonds, double bonds and triple bonds, using cocktail sticks and small polystyrene balls or locally available materials. Students should appreciate that the sharing of the pair of electrons in a covalent bond depends on the attraction which each atom has for the electrons. This is related to the electro negativity of the atom. Students should discuss the sharing of the pair of electrons in molecules between two atoms of the same element, or two atoms of similar electro negativity: • the electrons are attracted equally by the two atoms • the electron pair is shared equally Students should discuss the sharing of the pair of electrons in molecules between two atoms which are significantly different in electro negativity e.g. hydrogen and chlorine in hydrogen chloride. • the chlorine atom is more electronegative than the hydrogen atom • the chlorine atom attracts the electrons in the covalent bond more strongly • the pair of electrons in the covalent bond are positioned closer to the chlorine atom than the hydrogen atom • the chlorine end of the bond is slightly negative relative to the hydrogen end • the relative charge is shown using the symbols δ^+ and δ^- • the bond is polarised and shown as $H^{\delta+}$ $Cl^{\delta-}$ Students should be able to identify polar bonds in covalent compounds and mark their polarity. Students should discuss the structure or different covalent bonds and determine whether they are non-polar or polar, if if they are polar, to what extent. Students could discuss how polarity affects the

Competencies	***Contents***	***Suggested Activities***
		properties of covalent compounds.
• Define coordinate covalent (dative) bond	• Coordinate covalent bond (dative bond)	Students should appreciate that another type of covalent bond exists in which both of the electrons which form a covalent bond between atoms is donated by one of the atoms. This is a coordinate covalent bond or dative bond.
• Illustrate the formation of coordinate covalent bond using suitable examples		This could include the formation of ammonium ion(NH^+_4) from hydrogen ion (H^+) and a molecule of ammonia(NH_3)  Students should be able to show the formation of coordinate bonds in some familiar substances including: • hydronium ion, H_3O^+ • carbon monoxide, CO Students should understand that a non-bonding pair of electrons is referred to as a lone pair. Students should appreciate the role of lone pairs of electrons in coordinate bonding. Students could name the covalent compounds and draw Lewis structures to show the covalent bonding.
• Explain the general properties of covalent compounds	• General properties of covalent compounds	Students should discuss general properties of covalent compounds including: • liquids or gases, and some are solids. • low melting points and boiling points • do not conduct an electric current when molten or in aqueous solution • solubility in non-polar solvents • insolubility in polar solvents like water
• Investigate the properties of given samples of covalent compounds.		Students could be given samples of covalent compounds and asked to investigate their properties.

Competencies	*Contents*	*Suggested Activities*
• Explain the formation of metallic bond	**3.4 Metallic bonding** *(1 period)* • Formation of metallic bond	Students should appreciate that metals consist of a matrix of positively charged ions in a sea of delocalised electrons. Students could discuss why metals are able to lose valence electrons so easily to give this structure. Students should know that metallic bonding is the result of electrostatic attraction between the positively charged metal ions and the negatively charged delocalised electrons. Students could discuss how the strength of the metallic bond is related to the atomic radius of the metal atom and the number of valence electrons which are delocalised. Metallic bonding is strongest where metals have small atomic radii and lose a large number of valence electrons. Students should appreciate that the ability of metals to conduct both heat and electricity is related to the movement of delocalised electrons. These are able to transfer energy and charge through the metal.
• Explain the electrical and thermal conductivity of metals in relation to metallic bonding. • Make a model to demonstrate metallic bonding.	• Properties of metallic bond	Students could discuss how some of the other properties of metals, such as malleability and ductility are related to metallic bonding. Students could demonstrate the difference in strength of metals using locally available metals. Students could make a model to demonstrate metallic bonding using small and large balls made of local materials.
• Define inter molecular force • Explain hydrogen bonding	**3.5. Intermolecular forces** *(4 periods)* • Hydrogen bonding	Students should know that intermolecular force is a force of attraction between molecules in a substance. Students should appreciate that: • a hydrogen bond is an intermolecular force of attraction between a highly polar hydrogen atom on one molecule and a highly electronegative atom on another • hydrogen bonding is the strongest of the intermolecular forces

Competencies	***Contents***	***Suggested Activities***
		Students should appreciate that: • the O-H bonds in water are polar due to the high electro negativity of oxygen • the bonds can be presented as $H^{\delta+}$- 0^{8-} • water molecules are attracted to each other $H^{\delta+}$- 0^{8-} ... H^{8+} - 0^{8-} • this attraction gives rise to hydrogen bonding
• Explain the effects of hydrogen bond on the properties of substances		Students should discuss the effects of hydrogen bonding on the properties of water including: • high melting point and boiling point • surface tension • formation of a meniscus • capillarity • coolant (in radiators)
• Describe Vander waals forces	• Van der Waals' forces	Students could discuss the effects of hydrogen bonding on other compounds. Students should appreciate that here are always forces of attraction between covalent molecules even if they are polar or not. Van der Waals' forces are weak electrostatic forces that bind both polar and non-polar molecules. They may be: • dipole-dipole forces • induced dipole-induced dipole or dispersion forces
• Explain dipole-dipole force • Give examples of molecules with dipole-dipole forces.	- dipole-dipole forces	Students should understand that a dipole which is two points of equal but opposite charges, separated by a distance. Students should appreciate that diploes exist in polar molecules and oppositely charged ends of dipoles on different molecules are attracted to each other. Students could be asked to give examples of molecules with dipole-dipole forces. This could include interaction between similar molecules like HCl or different molecules like HCl and water.
• Explain dispersion forces	- dispersion forces	Students should understand that dispersion

Competencies	Contents	Suggested Activities
• Give examples of molecules in which the dispersion force is important.		forces occur in non-polar molecules. They are the result of temporary fluctuations in electron density. Within any bond, a slight temporary movement of bonding electrons towards one or other of the atoms will cause an induced dipole. This will result in molecules being attracted to each other. Students could be asked to give examples of molecules which contain dispersion forces. This could include molecules like neon(Ne) and methane(CH_4). Students should also recognize that dispersion forces exist in all types of molecules.
• Compare and contrast the three types of inter molecular forces.		Students should identify molecules where there are dipole forces and where there are dispersion forces. Students should describe the similarities and the differences among the hydrogen bond, dipole- dipole force and dispersion (London) force.

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level
Students working at the minimum requirement level will be able to: Define chemical bonding and explain why it is formed between the atoms; explain the ions, ionic bonding and its formation; give examples of simple ionic compounds and draw their Lewis structures, explain the general properties of ionic compounds; investigate the properties of a given sample of ionic compounds; define covalent bonding and describe its formation; give examples of different types of covalent molecules and draw their Lewis structures; make models of covalent compounds; explain polarity of covalent compounds; distinguish between polar and non-polar covalent molecules define coordinate covalent bond and illustrate its formation; investigate properties of a given sample of covalent compounds, explain metallic bonding; explain electrical and thermal conductivity of metals; make model to demonstrate metallic bond, define intermolecular force; explain hydrogen bond and its effect on the properties of substances; describe Vander waals forces, explain dipole-dipole forces and give examples of molecules with dipole - dipole forces, explain dispersion forces and give examples of molecules with dispersion forces and compare and contrast the three types of intermolecular forces.

Students above minimum requirement level
Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level
Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day

Unit 4: Chemical reactions and stoichiometry *(37 periods)*

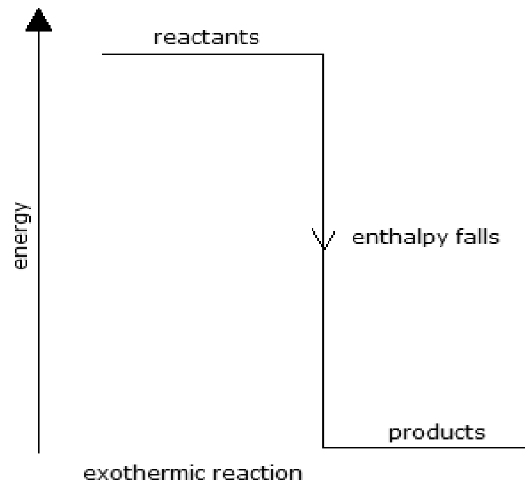
Unit outcomes: Students will be able to:

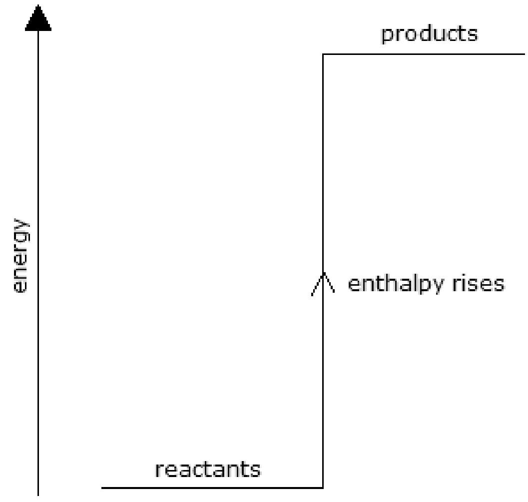
- Understand fundamental laws of chemical reactions and know how they are applied.
- Develop skills in writing and balancing chemical equations.
- Understand energy changes in chemical reactions.
- Know types of chemical reactions.
- Develop skills in solving problems based on chemical equations (mass - mass, volume - volume and mass - volume problems).
- Develop skills in determining the limiting reactant, theoretical yield, actual yield and percentage yield.
- Understand oxidation - reduction reactions and analyse redox reactions by specifying the oxidizing agent, the reducing agent, the substance reduced or oxidized.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**4. Chemical reactions and stoichiometry**	
• Define chemical reaction • Give some examples of chemical reactions	**4.1 Introduction** *(1 period)*	Students should appreciate that a chemical reaction is a change that takes place when one or more substances, called reactants, react alone or with each other to produce one or more new substances, called products. Reactants → Products Students should give different examples of chemical reactions. The students could also discuss the examples of changes brought about by the chemical reactions. These could include: • heating sugar • rusting of iron • Fermentation • Souring tella • Digestion of food
• State the law of conservation of mass and illustrate using examples • Demonstrate the law of conservation of mass using simple experiments	**4.2 Fundamental laws of chemical reactions** *(2 periods)* • The law of conservation of mass	Students should appreciate that in all types of chemical reactions mass is neither created nor destroyed. Students should be able to quote the law of conservation of mass: 'Matter cannot be created nor destroyed in a chemical reaction' Students could carry out an experiment to prove the law of conservation of mass. For example using the reaction between silver nitrate solution and dilute hydrochloric acid: $HCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + HNO_3$

Competencies	***Contents***	***Suggested Activities***
		• Place dilute hydrochloric acid in a conical flask to a depth of about 1 cm • Tie a thread of cotton around the top of a test tube • Half fill the test tube with silver nitrate solution • Place the test tube inside the conical flask so that it is held on a slant by the thread and place a bung in the top of the flask to hold the thread in place • Weigh the conical flask and contents • Tilt the flask so the silver nitrate solution pours into the dilute hydrochloric acid and a white precipitate of silver chloride is produced • Reweigh the conical flask and contents From their experiment students should show that the mass of the products is equal to the mass of the reactants. Students could also use the reaction between barium nitrate and sodium sulphate: $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$ • Make up solutions containing 2.61 g of barium nitrate and 1.42 g of sodium sulphate • Mix the solutions • Filter off the insoluble barium sulphate and dry it (2.33 g) • Evaporate the water from the filtrate and weigh the residue of sodium nitrate (1.70 g) Students should find that, within the limits of experimental error, the mass of the products equals the mass of the reactants.
• State the law of definite proportion and illustrate using examples • Demonstrate the law of definite proportion using a simple experiment. • State the law of multiple proportion and illustrate using examples.	• The law of definite (constant) proportion	Students should be able to quote the law of constant or definite proportion: 'The proportion by mass of each element in a pure compound is always the same however the compound is made' Students could carry out an experiment to prove the law of definite proportion. For

Competencies	*Contents*	*Suggested Activities*
	• The law of multiple proportion	example: • Make copper(II) oxide by heating copper powder • Make copper(II) oxide by the thermal decomposition of copper(II) carbonate • Take 1 g of each sample of copper(II) oxide • Reduce each sample of copper(II) oxide by heating in a stream of hydrogen • Weigh the copper that remains in each case Students should find that the mass of copper is the same in each case therefore the proportions of copper and oxygen in each sample of copper (II) oxide is the same.
• Describe the conventions used to write chemical equation • Balance chemical equations using inspection method • Balance chemical equations using the Least Common Multiple (LCM) method • Explain energy changes in chemical reactions	**4.3 Chemical equations** *(3 periods)* • Writing chemical equation • Balancing chemical equation **4.4 Energy changes in chemical reactions** *(3 periods)*	Students should appreciate that a chemical equation is a means of describing a chemical reaction: • Qualitatively – by identifying the reactants and products • Quantitatively – by identifying the relative proportion of each reactant and product Students should use the following steps to write an equation: 1. Write a word equation 2. Substitute the words by symbols and formulas 3. Balance the equation so that there are equal numbers of atoms of each type of element on each side of the equation Students should be able to balance an equation: • by inspection method • by the least common multiple method Students should have plenty of practice of writing chemical equations. As far as possible these should relate to reactions which are familiar to students. Examples should increase in difficulty as students become more able. Students should appreciate that changes in energy take place during a chemical reaction. Students should understand that enthalpy is a

Competencies	***Contents***	***Suggested Activities***
• Define endothermic reaction • Describe endothermic reaction	• Endothermic reaction	measure of the internal energy of a substance. It is represented by the symbol *H*. The energy change that takes place during a chemical reaction are the result of a Students should know that in an endothermic reaction: • Heat is taken in from the surroundings • The internal energy of the reactants is less than the internal energy of the products • There is a rise in enthalpy therefore the value of ΔH is positive change in enthalpy and is represented as ΔH.
• Define exothermic reaction • Describe exothermic reaction	• Exothermic reaction	Students should know that in an exothermic reaction: • Heat is given out to the surroundings • The internal energy of the reactants is more than the internal energy of the products • There is a fall in enthalpy therefore the value of ΔH is negative Students should appreciate that most of the reactions they will see and carry out in the laboratory will be exothermic.
• Illustrate endothermic and exothermic reactions using diagrams	• Energy diagrams for endothermic and exothermic reactions	Students should be able to draw energy level diagrams to represent exothermic and endothermic reactions. These have the general form:  $H_r > H_P$ ∴ ΔH is negative ($\Delta H < 0$)

Competencies	***Contents***	***Suggested Activities***
		H_P - heat content of products  endothermic reaction $H_P > H_r$ $\therefore$ ΔH is positive ($\Delta H > O$) Students should carry out some exothermic and endothermic reactions and understand that heat is given out or taken in either by direct observation or by measuring temperature rise with a thermometer. These could include: Exothermic • Burning gas • Adding water to anhydrous copper(II) sulphate • Adding a few drops of concentrated sulphuric acid to water • Mixing dilute hydrochloric acid and sodium hydroxide solution
• Conduct simple experiment to demonstrate exothermic and endothermic reactions		• Thermite process Endothermic • Adding solid ammonium nitrate to water • Adding solid potassium nitrate to water
• Describe the importance of chemical changes in the production of new substances and energy	• Importance of chemical changes	Students should appreciate the importance of the energy released in some exothermic reactions. These should include: • Combustion of fuels • Oxidation of glucose during cell respiration

Competencies	***Contents***	***Suggested Activities***
• List the four types of chemical reactions	**4.5 Types of chemical reactions** *(3 periods)*	Students should appreciate that chemical reactions can be classified into types or groups and that all of the reactions within a group have certain similarities.
• Define combination reaction and give examples • Conduct some experiments on combination reactions in groups.	• Combination	Students should know that combination or synthesis reactions involve the reaction of one element with another to form a compound. Students should discuss examples of combination reactions. These could include: • Sodium + chlorine = sodium chloride • Carbon + oxygen = carbon dioxide • Copper + oxygen = copper(II) oxide • Iron + sulphur = iron sulphide • Hydrogen + oxygen = water • Magnesium + oxygen = magnesium oxide Students should discuss the general features of a combination reaction. Students could carry out some combination reactions.
• Define decomposition reaction and give examples • Conduct some experiments on decomposition reactions in groups.	• Decompositi on	Students should know that decomposition reactions involve a reactant breaking down to produce two or more products. Students should discuss examples of decomposition reactions. These could include: • Thermal decomposition of Group 1 nitrates to nitrites and oxygen • Thermal decomposition of Group 2 and Transition Metal nitrates to oxides , oxygen and nitrogen dioxide • Thermal decomposition of Group 2 and Transition Metal carbonates to oxides and carbon dioxide Students should discuss the general features of a decomposition reaction. Students could carry out some decomposition reactions.
• Define single	• Single displacement	

Competencies	***Contents***	***Suggested Activities***
and give examples • Conduct some experiments on simple displacement reactions in groups.		reactive element displacing a less reactive element from a compound. Students should discuss examples of single displacement reactions. These could include: • Reactive metals with water – displacing hydrogen e.g. sodium + water • Reactive metals with dilute acids – displacing hydrogen e.g. zinc + dilute hydrochloric acid • Metal – metal ion reactions – atoms of a more reactive element displace ions of a less reactive metal from solution e.g. iron + copper(II) sulphate solution • Halogen – halide ion reactions – atoms of a more reactive halide displace ions of a less reactive halide from solution e.g. chloride + potassium bromide solution. Students should discuss the general features of a single displacement reaction. Students could carry out some single displacement reactions.
• Define double decomposition reaction and give examples. • Conduct some experiments on double displacement reactions in groups.	• Double decomposition	Students should know that double decomposition reactions involving reacting solutions of two soluble compounds to form two products, one of which is soluble and one of which is not. The products are separated by filtration. Students should discuss examples of double decomposition reactions. These could include: • Reactions to precipitate silver halides • Reactions to precipitate lead halides • Reactions to precipitate barium sulphate Students should discuss the general features of a double decomposition reaction. Students could carry out some double displacement reactions.

Competencies	***Contents***	***Suggested Activities***
• Deduce mole ratios from balanced chemical equations	**4.6 Stoichiometry** *(10 periods)* • Molar ratios in balanced chemical equations.	The student should revise the terms mole, motor mass and actual mass and related problems studied in grade 8 unit 5. Students should appreciate that stoichiometry is the study of the different amounts of substances which react to give new substances. Students should appreciate the significance of the different numbers of reactant and product particles in a balanced chemical equation. Students should appreciate that a balanced chemical equation indicates the molar quantities of reactants and products involved in a reaction. Students should be able to interpret a balanced chemical equation in terms of ratios of moles of the reactants and products. Students should interpret a variety of balanced equations in this way.
• Solve mass-mass problems based on the given chemical equation • Define molar volume	• Mass-mass relationships.	Students should be able to use balanced chemical equations to deduce the ratio of reacting masses and use this to calculate the actual masses. For example to find the mass of AgBr produced from 1 gm of KBr: $KBr(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgBr(s)$ 1 mole 1 mole 1 mole 1 mole 119 g 170 g 101 g 188 g • 119 g of KBr yields 188 g of AgBr • therefore 1.00 g of KBr yields $\frac{1.00 \times 188}{119}$ = 1.58 g of AgBr Students should carry out a number of similar calculations on reacting masses using equations with different mole ratios.
• State Avogadro's principle		Students should appreciate that one mole of gas particles occupies a volume of 22.4 litre at standard temperature and pressure (STP) 0 °C and 101.325 Kpa.

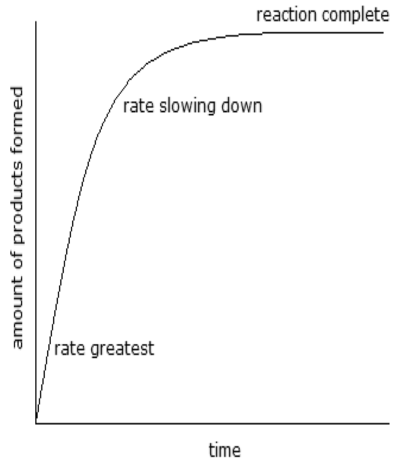
• Solve volume-volume problems based on the given chemical equation	• Volume-volume relationships.	Students should be aware of Avogadro's principle which states that: 'equal volumes of different gases under the same conditions of temperature and pressure contain equal numbers of particles; Students should be able to use balanced chemical equations to deduce the ratio of reacting volumes and use this to calculate the actual volumes. For example to find the volume of HCl produced from 100ml of H_2: $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ 1 mole 1 mole 2 moles 22.4 L 22.4 L 44.8 L • 22.4 L of H_2 yields 44.8 L of HCl • 100 ml of H_2 yields $\frac{0.1 \times 44.8}{22.4}$ = 0.2 L = 200 ml of HCl Students should carry out a number of similar calculations on reacting volumes using equations with different mole ratios.
• Solve mass-volume problems based on the given chemical equation	• Mass - volume relationships.	Students should combine their knowledge of reacting masses and molar volume to carry out calculations involving both mass and volume. For example to find the volume of CO_2 produced from 0.50gm of $CaCO_3$ $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ 1 mole 1 mole 1 mole 100 g 56 g 44 g or 22.4 L • 100 g of $CaCO_3$ yields 56 g of CaO • 0.50 g of $CaCO_3$ yields $\frac{0.5 \times 56}{100}$ = 0.28 g • 100 g of $CaCO_3$ yields 22.4 L of CO_2 • 0.50 g of $CaCO_3$ yields $\frac{0.5 \times 22.4}{100}$ = 0.112 L Students should carry out a number of similar calculations on reacting masses and volumes using equations with different mole ratios.
• Define limiting and excess reactants • Determine limiting and	• Limiting and excess reactants.	

Competencies	***Contents***	***Suggested Activities***
excess reactants of a given chemical reaction • Show that the amount of product of a chemical reaction is based on the limiting reactant		Students should appreciate that in a chemical reaction involving two reactants, the reaction will stop when all of one reactant has been used up no matter how much of the second reactant remains. Students should understand that when two reactants are not in the mole ratio in which they react then one will be the limiting reactant and the other will be in excess. Students should carry out calculations on chemical reactions in which there is a limiting reactant. For example: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ 1 mole 2 moles 1 mole 1 mole 5.00 g of zinc is reacted with 3.65g of HCl • 5.00 g of zinc is $\frac{5.00}{65.4}$ = 0.076 mol • 3.65g of hydrochloric acid contains 3.65/ 36.5 = 0.1 mol. • The maximum amount of zinc that will react with 0.1 mol dilute hydrochloric acid is 0.050 mol • The dilute hydrochloric acid is the limiting reactant • The zinc is in excess • The amount of zinc that will remain at the end of the reaction is 0.076 – 0.050 = 0.026 mol which is 65.4 x 0.026 = 1.70 g • The amount of zinc chloride formed will be 0.050 mol which is (65.4 + 2 x 35.5) x 0.05 = 6.82 g • The amount of hydrogen gas formed will be 0.050 mol which is 22.4 x 0.050 = 1.12 L at room temperature and standard pressure
• Define the terms theoretical yield, actual yield and percentage yield	• Theoretical, actual and percentage yields.	Students should carry out a number of similar calculations, identifying the limiting reactant and finding the maximum amount of products that can be formed. Students should appreciate that: • given the equation for a chemical reaction and the quantities of reactants used, it is possible to calculate the yield

Competencies	*Contents*	*Suggested Activities*
• Calculate the percentage yield of a chemical reaction from given information		of products that should be obtained – the theoretical yield • not all of the reactants in a chemical reaction may go to form a desired product; some may be used up in unwanted side reactions. The actual amount of product obtained is – the actual yield • the success of a reaction can be assessed by comparing the ratio of the actual yield to the theoretical yield to give – the percentage yield Students should be able to calculate the percentage yield of a chemical reaction from a given data using the equation: $\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$ Students should appreciate that: • percentage yield will be between 0 – 100% • the higher the percentage yield the more successful the reaction
• Define redox reactions • Define the terms oxidation and reduction in terms of electron transfer	**4.7 Oxidation and reduction reactions** *(5 periods)* • oxidation • reduction	Students should know that redox reaction is a reaction that involves transfer of electrons. Students should understand that the term redox' is used to describe reduction and oxidation reactions and that these reactions occur together. Students should be able to define oxidation as a loss of electrons and reduction as a gain of electrons. They should also understand that in a redox reaction: • the oxidised species loses one or more electrons • the reduced species gains one or more electrons Students should know that the oxidation number or oxidation state: • is the charge that the atom carries in its compounds

Competencies	***Contents***	***Suggested Activities***
• Define oxidation number (oxidation state) • State oxidation number rules. • Determine the oxidation number of an element in a given formula • Determine the oxidation number of an element in a given formula • Describe the oxidizing and reducing agents	• oxidation number • Oxidizing and reducing agents.	• refers to a single atom of the element • has both sign and numerical value Students should state oxidation number rules. Students should be able to deduce the oxidation number of elements in compounds from their formulas Students should do simple examples e.g. $MgCl_2$ • oxidation state of Mg = +2 • oxidation state of Cl = -1 And more complex examples where an atom is part of an ion e.g. SO_4^{2-} • oxidation state of O = -2 • total oxidation state of O=4 x -2 = -8 • overall charge on the ion = -2 • oxidation state of S = -2 – (-8) = +6 Students should appreciate that an oxidising agent is one which brings about oxidation but is itself reduced. Students should discuss examples of oxidising agents which they may have already met in chemical reactions such as: • chlorine • potassium manganate(VII) • potassium chromate(VI) • potassium dichromate(VI) • sodium chlorate(V) • manganese(IV) oxide Students should appreciate that a reducing agent is one which bring about reduction but is itself oxidised. Students should discuss examples of reducing agents which they may already have met in chemical reactions such as: • carbon • carbon monoxide • hydrogen • iron(II) salts • acidified iodide ions • metals such as sodium, magnesium and

Competencies	***Contents***	***Suggested Activities***
• Analyze a given redox reaction by specifying the substance reduced the substance oxidized, the oxidizing and reducing agents	• Analysing redox reactions	• zinc • sodium thiosulphate • sodium sulphite Students should be able to give examples of redox reaction and identify the oxidised and reduced species. They could start with simple reactions e.g. $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$ • oxidised species is Fe • reduced species is Cu^{2+} and progress to more difficult examples e.g. e.g. $SO_3^{2-}(aq) + H^+ (aq) + MnO_4^- (aq) \rightarrow SO_4^{2-} (aq) + H_2O(l) + Mn^{2+} (aq)$ • oxidised species is SO_3^{2-} (S oxidised from +4 to +6) • reduced species is MnO_4^- (Mn reduced from +7 to +2)
• Distinguish between redox and non-redox reactions	- Differences between redox and non - redox reactions	Students should appreciate that a chemical reaction can only be described as a redox reaction if one species is reduced and another species is oxidised. There are many reactions in which oxidation and reduction do not occur e.g. in double decomposition reactions: $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$ Students should find the oxidation states of the metal ions in both the reactants and products and satisfy themselves that no oxidation or reduction has taken place.
• Define a reaction rate	**4.8 Rate of chemical reaction and chemical equilibrium** *(10 periods)* • Reaction rate	Students should understand that the rate of a reaction is the rate at which reactants are converted to products. Students could discuss different ways in which the rate of a reaction can be monitored. These could include: • Change of colour • Volume of gas evolved • Amount of precipitate formed • Loss or gain of mass
• Describe a reaction rate using graphs		Students should be able to draw a graph showing how the rate of a chemical reaction changes over time and relate the rate to the slope of the graph.

Competencies	***Contents***	***Suggested Activities***
		
• Carry out an experiment to illustrate the relative rate of reactions		Students should carry out experiments to study how the rate changes during a chemical reaction. These could include: • Metal + dilute acid e.g. zinc + dilute sulphuric acid • Metal carbonate + dilute acid e.g. calcium carbonate + dilute hydrochloric acid In both of these types of reaction the rate can be monitored by: • Measuring the loss of mass over time • Measuring the volume of gas produced over time
• List the preconditions for a chemical reaction to occur • Explain how collision, activation energy and proper orientation of reactants cause a chemical reaction to occur	• Preconditions for a chemical reaction	Students should be able to list the preconditions for a chemical reaction to take occur – collision, activation energy and proper orientation
• List factors that affect rate of chemical reaction	• Factors affecting rate of chemical reaction	Students should appreciate that in order for a chemical reaction to take place the reacting particles must collide with each other with sufficient force – an effective collision. Students should discuss how providing reacting particles with activation energy: • Increases the kinetic energy of the particles • Increases the chances that collisions between particles will be effective

Competencies	***Contents***	***Suggested Activities***
• Explain the effect of changes in temperature, concentration or pressure and surface area on the rate of a chemical reaction • Explain the effect of catalysts on the rate of chemical reaction		Students should appreciate that the rate of a chemical reaction depends on: • Nature of reactants • Temperature • Concentration or pressure • Surface area • Use of a catalyst Students should discuss these factors in terms of collision theory. This could include: • Temperature – an increase in temperature increases the kinetic energy of the reacting particles so there is a greater chance of collisions being effective • Concentration or pressure – increasing the concentration (of reacting solutions) or pressure (of reacting gases) increases the chances of particles colliding with each other as there will be more particles per unit volume • Surface area – increasing the surface area of a solid reactant increases the chances of collisions with the other reactant • Use of a catalyst – a catalyst reduces the activation energy so particles collisions can be effective at lower temperatures
• Carry out an activity on how the factors affect the rate of chemical reaction		Students could investigate the effects of each of these factors experimentally e.g. • Temperature – metal-acid reaction using acid of different temperatures • Concentration – metal-acid reaction or metal carbonate-acid reaction using acid of different concentrations • Surface area – zinc-acid reaction using granulated zinc and zinc powder; calcium carbonate-dilute hydrochloric acid using chippings and powdered calcium carbonate • Catalyst – decomposition of hydrogen peroxide using manganese(IV) oxide
• Define the terms reversible reaction and irreversible reaction.	• Reversible and irreversible reactions	Students should appreciate that some chemical reactions are difficult or impossible

• Define chemical • equilibrium • Describe the characteristics of chemical equilibrium	• Chemical equilibrium	to reverse whereas others can be reversed. Students should discuss examples of irreversible and reversible reactions. Students should understand that where a reaction is reversible, as soon as any product is formed, the reverse reaction will begin unless the products are removed. Forward reaction Reactants ⇌ Products Reverse or backward reaction Students should appreciate that: • The symbol '$\rightleftharpoons$' is used to show that a reaction is reversible • At some point in time the rate of the forward reaction becomes equal to the rate of the backward reaction • At this point the concentration of each reactant and each product remains constant • The reaction is said to be in equilibrium • This is described as dynamic equilibrium since the reactions are still taking place even though the concentrations of the reactants and products remain constant.
• Write the expression for equilibrium constant of a reversible reaction	• Equilibrium constant	Students should be aware of the conventional use of square brackets [] to show the concentration of a species in a reversible reaction.
• State Le Châtelier's principle • Use Le Châtelier's principle to explain the effect of changes in temperature, pressure and concentration of reaction at equilibrium	• Factors that affect chemical equilibrium	Students should know that for a reversible reaction represented by: $aA + bB \rightleftharpoons cC + dD$ • The ratio of the equilibrium concentrations of products raised to the power of their coefficient in a balanced equation is constant at a given temperature • This is called the equilibrium constant and has the symbol K_c • $K_c = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$ Students should carry out calculations using given data to find the equilibrium constant for reaction at equilibrium.

Competencies	***Contents***	***Suggested Activities***
		Students should be able to state Le Châtelier's principle: 'If a condition is changed, the position of equilibrium will move in such a way as to oppose the change and restore the original equilibrium conditions' Students should discuss how this principle accounts for the effects of changes in temperature and pressure (concentration) in reactions including: • $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ • $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ And in industrial processes including: • The Haber process $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ • The Contact process $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: Define chemical reactions and give examples; state and demonstrate the law of conservation of mass and the law of definite proportions, write chemical equations and balance them using inspection and LCM methods; explain energy changes in chemical reactions, conduct simple experiments on endothermic and exothermic reactions; describe the importance of chemical changes; define combination, decomposition, single displacement, double displacement reactions and give examples for each; conduct experiments on combination; decomposition, simple displacement and double displacement reactions; define molar volume, limiting reactant, excess reactant, theoretical yield, actual yield and percentage yield, deduce molar ratios from balanced chemical equations and solve problems on mass - mass, volume - volume and mass - volume relationships, determine the limiting reactant of a chemical reaction and show that the amount of product is based on the limiting reactants; calculate the percentage yield of a chemical reaction; define redox reactions, oxidation ,reduction and oxidation number, state oxidation number rules and determine the oxidation number of an element; describe the oxidizing and the reducing agents, distinguish between redox and non -redox reactions; analyze a given redox reaction; define and describe a reaction rate; list and explain the preconditions for a chemical reaction to occur; list and explain factors that affect the rate of a chemical reaction; carryout an activity of factors that affect the rate of a reaction; define the terms reversible reactions, irreversible reactions and chemical equilibrium; describe the characteristics of chemical equilibrium and write an expression for its equilibrium constant, state Lechattlers principle and use it to explain the effect of the factors that affect chemical equilibrium.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.

Unit 5: The physical states of matter *(20 periods)*

Unit outcomes: Students will be able to:

- Understand the kinetic molecular theory and properties of the three physical states of matter.
- Know the behaviour of gases by using the variables volume, temperature, pressure and number of moles.
- Know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, molar heat of solidification.
- Understand gas laws.
- Develop skills in solving problems to which the gas laws apply.
- Perform activities to illustrate gas laws.
- Carryout experiments to determine the boiling points of liquids and the melting point of solids.
- Demonstrate an experiment to show phase changes.
- Demonstrate scientific inquiry skills: observing, predicting, comparing and contrasting, measuring, interpreting data, drawing conclusion, applying concepts and making generalizations.

Competencies	***Contents***	***Suggested Activities***
Students will be able to:	**5. The physical states of matter**	
• Name the three physical states of matter. • Give examples for each of the three physical states of matter • State kinetic theory of matter	**5.1 Introduction** *(1 period)* **5.2 Kinetic theory and properties of matter** *(3 periods)*	Students should know that there are three physical states of matter: solids, liquids and gases. Students should also be aware that, although it is seldom found in everyday life, there is the fourth state known as the plasma state. Students should give everyday examples of substances that occurin each of these states.
	• Kinetic theory of matter	Students should know the kinetic theory of matter: • All matter is composed of particles which are in constant motion • The particles possess kinetic energy (movement energy) and potential energy • The difference between the three states of matter are due to their energy contents and the motion of the particles
• Explain the properties of the three physical states of matter in terms of kinetic theory	• Properties of matter	Students should know that in a solid: • Particles are in fixed positions • Particles are able to vibrate but not translate

Competencies	***Contents***	***Suggested Activities***
• Compare and contrast the three physical states of matter		Students should discuss how the kinetic theory explains the properties of solids e.g. fixed shape, can't be compressed Students should know that in a liquid: • The particles have more energy than in a solid • The particles are not in fixed positions but able to move • The particles are in continual constant motion • The particles are closely packed but less so than in a solid Students should discuss how the kinetic theory explains the properties of liquids e.g. taking the shape of a container, can only be compressed a small amount. Students should know that in a gas: • The particles have more energy than in a liquid • There are large distances between particles compared to liquids and solids • Particles move about in all directions at high speed Students should discuss how the kinetic theory explains the properties of gases e.g. diffusion to fill the space available, easy to compress Students should discuss the differences between the properties of the three states of matter based on the kinetic theory.
• Explain the assumptions of kinetic molecular theory of gases • Describe properties of gases using kinetic molecular theory • Describe the behaviour of gases by using the variables V (volume),T	**5.3 The gaseous state** ***(11 periods)*** • ***The kinetic molecular theory gases*** • The gas laws	Students should state the kinetic molecular theory assumptions and use them in explaining properties of gases Students should appreciate that the behaviour of gases is explained by a series of gas laws which study the relationships among variables: • V denotes volume • T denotes temperature • P denotes pressure • n denotes number of moles Students should also know some

Competencies	***Contents***	***Suggested Activities***
(Temperature), P (pressure) and n (number of moles)		common units we use for the above mentioned variables
• State Boyle's law • Perform an activity to show changes in volume and pressure of gases to illustrate Boyle's law • Apply Boyle's law in solving problems	- Boyle's law	Students should be able to state Boyle's law: 'The volume of a given mass of gas is inversely proportional to its pressure provided that the temperature remains constant' Students should understand that this can be expressed mathematically as: • $P \propto \frac{1}{V}$ or PV = K where K is a constant When we change the pressure or volume at constant temperature it follows that: • $P_1V_1 = P_2V_2$ where subscripts denote initial and final conditions Students should carry out a practical activity to prove the validity of Boyle's law as follows: • Two tubes partially filled with mercury joined by rubber tubing giving a U-tube arrangement • One tube is calibrated, contains air and is sealed by a tap • By moving the un-calibrated tube up and down the volume of air in the calibrated tube can be varied • The pressure exerted on the air is obtained from the difference in height of the mercury in the two tubes • Students should plot a graph of P against 1/V and obtain a straight line Students should carry out calculations using Boyle's law. Students should be able to state Charles' law: 'the volume of a fixed mass of gas at constant pressure is directly proportional to its absolute temperature' Students should understand that this can be expressed mathematically as:
• State Charles' law	- Charles' law	• $V \propto T$ or $\underline{V}$ = K where K is a

Competencies	Contents	Suggested Activities
• Perform an activity to show changes in volume and temperature of gases to illustrate Charles' law. • Apply Charles' law in solving problems		constant T When we change volume or temperature at constant pressure it follows that: • $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ where subscripts denote initial and final conditions Students should carry out a practical activity to demonstrate Charles' law as follows: • Set up a round-bottomed flask fitted with a bung and delivery tube • Invert the flask and place the delivery tube in a beaker of water • Investigate what happens to the volume of air in the flask when it is heated or cooled by observing whether bubbles of air are forced out or water is drawn up the delivery tube Students should carry out calculations using Charles' law.
• Derive combined gas law equation from Boyle's law and Charles' law • Use the combined gas law to calculate changes in volume, pressure or temperature	- Combined gas law	Students should appreciate that it is possible to combine Boyle's law and Charles' law to give one law called the combined gas law. Students should understand that this can be expressed mathematically as: • $\frac{PV}{T} = K$ where K is a constant More generally for changing conditions: • $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ Students should carry out calculations using the combined gas law.
• Define an ideal gas • Derive an ideal gas equation from Boyle's	- Ideal gas equation	Students should appreciate that an ideal gas is a hypothetic gas that exactly obeys the gas laws. In reality real gases only obey the gas laws closely at high temperatures and low pressures as under these conditions their particles are far apart. Students should understand that the ideal gas law is a combination of Boyle's law, Charles' law and

Competencies	***Contents***	***Suggested Activities***
law, Charles' law and Avogadro's law • Compare the nature of real gases with ideal gases • Solve problems related to ideal gas equation		Avogadro's law. (Grade 9 unit 4). • Boyle's law $V \propto 1/P$ (n and T constant) • Charles' law $V \propto T$ (n and P constant) • Avogadro's law $V \propto n$ (P and T constant) It follows from these that: • $V \propto \frac{nT}{P}$ • $V = \frac{RTn}{P}$ or $PV = nRT$ • where R = molar gas constant = 8.314 $\frac{L.Kpa}{mol.k}$ or 0.082 $\frac{L.atm}{mol.k}$ or 8.314 $\frac{J}{mol.k}$ Students should carry out calculations using the ideal gas law.
• Define diffusion • State Graham's law of diffusion • Carry out an activity to compare the rate of diffusion of two different gases • Apply Graham's law of diffusion in solving problems	- Graham's law of diffusion	Students should understand that diffusion of a gas is the spreading out of its particles so that they are evenly distributed through the whole of the space available. Students should know Graham's law of diffusion: 'At constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportionally to the square root of its density, ρ, or molar mass, M.' Students should understand that this can be expressed mathematically as follows: • $r \propto \frac{1}{\sqrt{\rho}}$ or $r = \frac{K}{\sqrt{\rho}}$ where k is a constant • $r \propto \frac{1}{\sqrt{m}}$ or $r = \frac{K}{\sqrt{m}}$ where k is a constant √M Also the rate at which a gas diffuse is inversely proportional to the time taken therefore:

Competencies	Contents	Suggested Activities
		• $r \propto \frac{1}{t}$ so comparing two gases $\frac{r_1}{r_2} = \frac{t_2}{t_1}$ It also follows that • $\frac{r_1}{r_2} = \sqrt{(\rho_2 / \rho_1)} = \sqrt{(m_2/m_1)}$ Students should compare the rate of diffusion of gases as follows: • Porous pot fitted with a cork and delivery tube • Delivery tube attached to a U-tube containing coloured water • Porous pot placed in beaker of hydrogen • Pressure in pot increases as hydrogen diffuses in quicker than the air diffuses out • Porous pot placed in beaker of carbon dioxide • Pressure in pot decreases as air diffuses out quicker than carbon dioxide diffuses in • Pressure changes shown by levels of coloured water in the limbs of the U-tube Students should carry out calculations using the different versions of Graham's law
• Explain the terms: evaporation, boiling condensation, vapour pressure, boiling point, molar heat of vaporization and molar heat of condensation	**5.4 The liquid state** *(3 periods)*	Students should recall the simple properties of liquids from the work carried out at the start of this unit. They should be able to account for these properties using the kinetic theory of matter. Students should be able to explain a variety of terms relating to the liquid state including:
	• Boiling	• Boiling – change from a liquid to a gas at a definite temperature, the boiling point of the liquid
	• Condensation • Energy changes in	• Condensation – change from a gas to a liquid at the boiling point of the liquid

Competencies	***Contents***	***Suggested Activities***
	boiling and condensation	• Molar heat of vaporisation – the amount of heat needed to convert 1 mole of a liquid at its boiling point to a gas (the same amount of energy that is released when 1 mole of vapour at the condensing point of a vapour condenses to become liquid) • Molar heat of condensation - the amount of heat realised when 1 mole of a gas is converted to a liquid at its condensation point
	• Evaporation	• Evaporation – change of a liquid to a gas at any temperature. Students should discuss the differences between boiling and evaporation:
	• Vapour pressure	• boiling is at a fixed temperature and occurs throughout a liquid • Evaporation is at any temperature and occurs only at the surface of a liquid
• Carry out an activity to demonstrate the concept of vapour pressure • Carry out an activity to determine the boiling points of water and ethanol	• Determination of vapour pressure and boiling point of a liquid	Students should be able to explain vapour pressure and appreciate that a liquid boils when its saturated vapour pressure becomes equal to atmospheric pressure Students should carry out an experiment to demonstrate vapour pressure by connecting a round-bottomed flask with bung and delivery tube, containing some liquid, to a manometer. As the liquid is warmed the vapour pressure increases and this is detected by the manometer. Students could quantify the above experiment and collect data which could then be analysed. Students should carry out an experiment to measure the boiling points of water and ethanol just above the surface of the liquids as they boil. Students could discuss why the thermometer y should not be placed in the boiling liquid – superheating.

Competencies	*Contents*	*Suggested Activities*
• Explain the terms melting, fusion, sublimation, melting point, freezing point, molar heat of fusion and molar heat of solidification.	**5.5 The solid state** *(2 periods)*	Students should recall the simple properties of solids from the work carried out at the start of this unit. They should be able to account for these properties using the kinetic theory of matter. Students should be able to explain a variety of terms relating to the solid state including:
	• Melting	• Melting (fusion) – change from a solid to a liquid at a definite temperature, the melting point of the solid
	• Freezing	• Freezing (solidification) – change from a liquid to a solid at the freezing point of the liquid
• Describe phase changes.	• Energy changes in melting and freezing	• Molar heat of fusion – the amount of heat needed to convert 1 mole of a solid at its melting point to a liquid (the same amount of energy that is released when 1 mole of liquid at the freezing point of a liquid freezes to become solid
• Explain temperature changes associated with phase changes		• Molar heat of solidification - the amount of heat released when 1 mole of a liquid is converted to a solid at its freezing points.
	• Sublimation	• Sublimation – change directly from solid to gas without going through the liquid phase or the reverse i.e. the direct change from a gas to solid
• Determine melting point of ice.	• Determination of melting point.	Students could measure the melting point of ice by placing a thermometer in a funnel containing melting ice, suspended about a beaker. Students should round off the unit by discussing phase changes.
• Demonstrate an experiment to show the phase changes from ice to liquid water and then to water vapour.	• Phase changes	Students should carry out experiments in which they: • Convert ice to water • Convert water to steam • Convert steam to water • Convert water to ice Students should be able to name the

Competencies	*Contents*	*Suggested Activities*
		processes involved and state which are exothermic (give out heat), and which are endothermic (take in heat). Students should discuss some of the practical implications of phase changes e.g. why does a burn from steam at 100 °C more damaging to the skin than a burn from water at 100 °C?

Assessment

The teacher should assess each student's work continuously over the whole unit and compare it with the following description, based on the Competencies, to determine whether the student has achieved the minimum required level.

Students at minimum requirement level

Students working at the minimum requirement level will be able to: Name the three physical states of matter and give examples for each; state kinetic theory of matter; explain the properties of three physical states of matter; compare and contrast the three physical states of matter; describe the behaviour of gases by using the variables volume (V), temperature (T), pressure (P) and number of moles (n); state Boyle's law, Charles' law and Graham's law of diffusion; Perform activities to demonstrate Boyle's law and Charles' law; apply the gas laws in solving problems; derive combined gas law equation and ideal gas law equation; define ideal gas and diffusion; compare the nature of real gases with ideal gases; carryout an activity to compare the rate of diffusion of different gases; Explain the terms: evaporation, boiling, condensation, vapour pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, molar heat of solidification, carryout activities to demonstrate the concept of vapour pressure and to determine the boiling points of water, ethanol and the melting point of ice; describe phase changes and explain temperature changes associated with them and demonstrate an experiment to show the phase changes from ice to liquid water and then to water vapour.

Students above minimum requirement level

Students working above the minimum requirement level should be praised and their achievements recognized. They should be encouraged to continue working hard and not become complacent.

Students below minimum requirement level

Students working below the minimum requirement level will require extra help if they are to catch up with the rest of the class. They should be given extra attention in class and additional lesson time during breaks or at the end of the day.