PREFACE

Teaching higher order thinking skills is currently at the centre of educational attention. In general, measures of higher order thinking include all intellectual tasks that call for more than the retrieval of information. Six fundamental higher order thinking skills have been identified in the Chemistry Syllabus. They are problem solving skills, inquiring skills, reasoning skills, communicating skills, conceptualising skills and creative and innovative skills. These skills, together with intertwining ways of learning chemistry, thinking and using chemical knowledge are considered, important in chemistry education.

In this chemistry course, students will be working with their teacher and other students to develop the basic knowledge and skills which will help them understand chemistry more and to apply them in their daily life. They will learn how to show an interest in the creativity and innovation found in chemistry. In some lessons, students will participate in group activities to develop skills in scientific methods of investigation based on the concepts, theory, terms, facts, laws and principles related to main themes.

Skill Development

After learning this course, students will develop and practise higher order thinking skills: comprehension, analysis, synthesis and evaluation. They will be able to participate actively in all lessons through the **5 C's** as important **21st century skills for learning**:

- ✓ **Collaboration** – in lessons students will be working in groups, to share ideas with their classmates and to find the solution together.
- ✓ **Communication** – students will develop verbal and non-verbal communication skills in group works.
- ✓ **Critical Thinking and Problem Solving** – students will be given interesting problems to solve – finding and explaining solutions, looking for correcting errors.
- ✓ **Creativity and Innovation** – thinking 'outside the box' is an important 21st century skill. Students will be encouraged to explore new ideas and solve problems in new ways.
- ✓ **Citizenship** – students will join the school community and develop fairness and conflict resolution skills.

Important Features of This Textbook

- ♦ The High School Chemistry Curriculum covers **six main themes**: Particulate Nature of Substances, Periodicity, Chemical Calculations, Chemistry of Reactions, The Environment, and Organic Chemistry.
- ♦ There are **eight chapters** included in this Textbook:
  - Chapter 1 Chemistry : The Central Science;

- Chapter 2 Matter and Solutions;
- Chapter 3 The Electronic Structures of Atoms and Periodic Table;
- Chapter 4 The Quantities of Substances : Chemical Calculations;
- Chapter 5 Non-metals : Carbon, Oxygen and Halogens;
- Chapter 6 Acids, Bases and Salts;
- Chapter 7 Air, Water and Soil;
- Chapter 8 Fuels and Crude Oils.

◆ Each chapter starts with the ***introduction*** to the topic, containing an example of how the material covered in the chapter, followed by the **Learning Outcomes** of the chapter.

◆ In each section of the chapter, the text and illustrations describe and explain all of the facts and concepts that students need to know. **Review Questions** after each section give students a chance to check that they have understood the topic they have just read. There is a summary of **Key Terms** at the end of each section. Key Terms are highlighted in the text when they are first introduced. All end-of-chapter **Exercises** are designed to ensure that students have grasped major concepts, in addition to testing their understanding of the materials covered in the chapter. At the end of each chapter, the **Chapter Review** (a concept link) points out the summary and highlights of the chapter.

◆ In addition, the **"Chemistry in Society"** or **"Chemistry in Daily Life"** highlighted in colour, is introduced in each section of the chapter, reminding students that chemistry is so central and so intimately involved in almost every aspect of our material world. It introduces students to the important chemicals and substances that form the basis for the high standard of living and modern technology that they now enjoy. *It is suggested that teachers should not emphasise this section as exam-oriented teaching.*

◆ The key terms mentioned in all chapters are listed in the **Glossary** at the end of the Textbook.

Goals of Grade 10 Chemistry Textbook

The Grade 10 Chemistry Textbook has been written for two goals. The first goal is to teach the fundamental of chemical concepts, and the second goal is to teach basic critical thinking skills. Students will solve the problems and predict events. The overall goal is to produce a text that introduces the students to the relevance and excitement of chemistry. We hope that students will understand the benefits and hazards of the Material World through the knowledge they have learned.

CHAPTER 1

CHEMISTRY: THE CENTRAL SCIENCE

1.1 CHEMISTRY AS CENTRAL SCIENCE

Chemistry is an area of knowledge remarkable for its breadth and depth. Knowledge of chemistry is essential to improve the quality of our lives. For instance, faster electronic devices, stronger plastics, and more effective medicines and vaccines all rely on the innovations of chemists throughout the world. We cannot truly understand or even know very much about the world we live in or about our own bodies without knowing the fundamental concepts of chemistry.

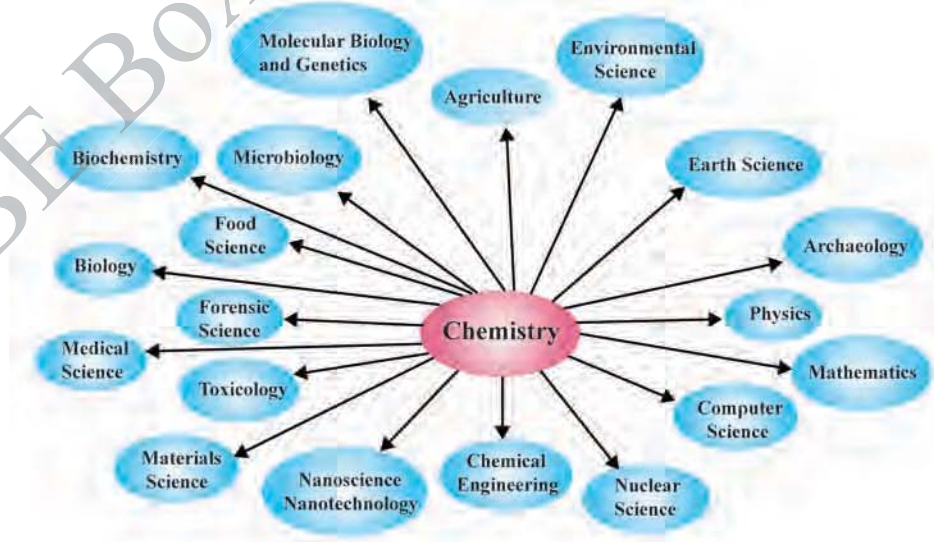
Climate change, water contamination, air pollution, food shortages and other societal issues are regularly featured in the media. However, did you know that chemistry plays a crucial role in addressing these challenges? As the 'Central Science', chemistry is woven into the fabric of practically every issue that our society faces today.

Learning Outcomes

After completing this chapter, students will be able to:

- describe the various branches of science and how society is impacted by them;
- recognise the role and impact of chemistry in daily life;
- discuss the importance of chemistry in daily life;
- distinguish and compare the branches of chemistry;
- develop and practice higher order thinking skills such as reasoning, analysis, synthesis, and evaluation.

Chemistry plays a central role in science and is often intertwined with other branches of science.



Just what is chemistry anyway? The usual definition is that chemistry is a study of matter and the changes it undergoes. What is matter? It is anything that has mass and occupies space. We change matter to make it more useful. Some matter we change to extract a part of its energy; for example, we burn gasoline to get energy to propel our automobiles. We practice chemistry everyday.

We practice chemistry when we cook in the kitchen, when we clean our house or paint our room, when we apply cosmetics, and when we take medicine or treat an injury. Our body takes oxygen from the air and combines it with part of the food we eat to provide us with energy for every activity we undertake. These are just a few of the ways in which chemistry impacts our daily lives.

So what is chemistry? It is a science that touches our life every moment. It deals with matter from the tiniest parts of atoms to the complex human body. It goes beyond the individual to affect society as a whole, shaping our civilisation.

1.2 MILESTONES IN THE HISTORY OF CHEMISTRY

Chemistry as a discipline has been around for a long time. The **history** of chemistry is an interesting and challenging one. In fact, chemistry is known to date back to as far as the prehistoric times. There are four categories that classified chemistry on the time line.

- Prehistoric time to the beginning of Christian era (about 300 BC) is classified as **black magic** period. The ancients believed the world made of 'four elements', which are earth, air, water and fire. The Greek philosophers were perhaps the first to formulate theories explaining the behaviour of matter without confirming their theories by experimentation. However, their philosophic point of view of nature, which can be attributed mainly to Aristotle, dominated the Greek culture. They used fire to bring about chemical changes. Examples include: extraction of metals from ores, making pottery and glazes, fermenting beer and wine, extracting chemicals from plants for medicine and perfume, making of fat into soap, manufacture of glass, and making of alloys like bronze and so on. These things and many others were accomplished without an understanding of the scientific principles involved.
- From about 300 BC to the end of 17th century, the experimental roots of chemistry are planted in **alchemy**, a mystical chemistry that flourished in China and Europe. Alchemists made several attempts to turn cheaper metals to gold, using the substance called the Philosopher's Stone. They also wanted to find an elixir that would enable people to live longer and cure all ailments. Alchemists never achieved these goals, but they discovered many new chemical substances and techniques such as distillation and extraction that are still used today. In 1661, Robert Boyle (1627-1691) developed the basic ideas about the behaviour of gases. His research progress was made in putting chemistry on a basic foundation.
- The field of chemistry began to develop rapidly in the 1700's. **Traditional chemistry** period started by the end of 17th century up to the middle of 19th century. In 1774 Joseph Priestley discovered a gas, later named oxygen. The chemistry was introduced to the science in 1768 by French chemist Antoine Lavoisier (1743–1794), who explained the law of conservation of matter based on the experimentation method. From this onward, scientists use

the experimentation method, also known as scientific method, rather than logical and theoretical method of the ancients. In 1803 John Dalton postulated Atomic Theory, which states that all matter is composed of atoms, which are small and indivisible. Amedeo Avogadro (1776 - 1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas, which we use today as Avogadro's constant.

- **Modern chemistry** starts from the middle of 19th century to the present time. The beginnings of modern chemistry were coming with the emergence of the experimental method when the works of scientists were characterised by a reliance on experimentation. Scientific facts remain the same, no matter who does the measuring. These facts are verified by repeated testing.

Few of the areas that have emerged as being especially important in modern chemistry are Synthesis, Separation techniques, Identification and assay, Materials, Polymers, Nanochemistry, Biochemistry, Molecular biology, Green chemistry, and Combinatorial chemistry.

Much of 20th century technology has grown out of scientific discoveries from radioactivity to artificial intelligence. Technological developments are used by scientists as tools for more discoveries. These developments in science and technology together with innovations are the basic current roots of the changing modern world.

1.3 IMPORTANCE OF CHEMISTRY

Chemistry is **important** because everything you do is **chemistry**. **Chemical** reactions occur when you breathe, eat, or just sit there reading. You are surrounded by materials and substances, all chemicals. Even your body is made of chemicals. The air you are breathing is a mixture of elements like oxygen and nitrogen. The book you are reading is made from wood pulp or cellulose which has been bleached and treated with various chemicals. The clothes you are wearing are probably made from synthetic chemicals called polymers, such as nylon or terylene. The seat you are sitting on is perhaps a plastic polymer, with polyurethane; foam seat padding and metal support. The room you are in is made from cement, plastics, concrete and glass, all of which are chemicals. Chemicals provide us with luxuries and improve our leisure time.

Some chemicals are toxic. Some causes cancer. Some chemicals are also beneficial. Some can save lives. Many are useful. All matter is made of chemicals, so the **importance** of chemistry is that it is the study of everything. **Chemistry deals with everything**. Perhaps a better understanding of chemistry would enable us to control the uses of chemicals so that we could maximise their benefits and minimise the risk involved in their use.

1.4 BRANCHES OF CHEMISTRY

There are many branches of chemistry or chemistry disciplines. The different branches focus on different aspects of matter. The five main **branches** are considered to be **Organic chemistry, Inorganic chemistry, Physical chemistry, Analytical chemistry,** and **Biochemistry (**Table 1.1**)**. In addition, Nuclear chemistry, Environmental chemistry, Industrial chemistry, Polymer chemistry, Materials chemistry, Nanochemistry, Green chemistry, Agricultural chemistry, Theoretical chemistry, etc., are other branches of chemistry.

Table 1.1 Branches of Chemistry

Branch	Areas of emphasis	Examples
organic chemistry	most carbon-containing chemicals which are hydrocarbons and their derivatives	pharmaceutical, plastics
inorganic chemistry	in general, matter that does not deal with hydrocarbons	minerals, metals and non-metals, semi-conductors
physical chemistry	the behaviour and changes of matter and the related energy changes	reaction rates, reaction mechanisms
analytical chemistry	components and compositions of substances	food nutrients, quality control (QC) and quality assurance (QA)
biochemistry	matter and processes of living organisms	metabolism, fermentation

1.5 UNDERSTANDING CHEMISTRY

By studying and understanding chemistry, you will become a global citizen in the twenty first century where you have to live and act more intellectually, mature and with greater understanding and satisfaction in an increasingly complex civilisation. To justify why we have to study chemistry, there are several reasons:

- Chemistry is clearly very broad. It is so important to the future of life on this planet that every educated citizen should have some knowledge of its scientific basis.
- The principles of chemistry are needed to understand the nature of every form of matter. For example, metals, drugs, gasoline, foods, the earth's crust, water, atmosphere, radioactive materials and the human brain cell, all have properties determined mainly by chemical principles.
- Furthermore, a course in chemistry can be a fascinating experience, because it helps you to understand yourselves and your surroundings in everyday living. Consumer aspects of chemistry and the chemistry of common things can also be known from the fundamental ideas of chemistry.
- Agriculture also uses chemistry in many areas. Chemistry is very important for the food and beverage industry and medical industry. Chemistry has helped in establishing industries which manufacture utility goods, such as acids, alkalis, soaps, detergents, dyes, polymers, metals, etc. These industries contribute in a big way to the economy of a nation and generate employment.

- Chemistry is also required for many fields of study. For example, courses in Pharmacy and Medicine require applicants to have knowledge of chemistry; cutting-edge of today's digital technology requires knowledge on the principles of chemistry.

It is for these reasons, the fact that chemistry is everywhere and does affect all aspects of our lives, that it is necessary to study it. Understanding basic chemistry and chemical terms will help to make your material world more meaningful.

Chemistry is considered as an experimental science. You should recognise that the principles and laws of nature are the results of extensive observations and speculative analyses refined over many investigations. Basic process skills such as observing, classifying, inferring, communicating (through diagram, graph, chart, etc.), measuring, predicting and using numbers will be developed if you study chemistry.

You are expected to enhance the development of these skills and use them to construct your chemical knowledge, and hence engage in life-long learning.

1.6 THE PRINCIPAL GOALS IN BASIC EDUCATION HIGH SCHOOL CHEMISTRY

Basic chemistry is the branch of science that studies the preparation, properties, structures and reactions of material substances. Chemistry contributes to a large extent in the development and growth of a nation. A developing country, like Myanmar, needs talented and creative chemists. To be a good chemist, one needs to understand the basic concepts of chemistry. On this context, there are **six main themes** in Basic Education High School Chemistry Course: **Particulate nature of substances, Periodicity, Chemical calculations, Chemistry of reactions, the Environment and Organic chemistry**. Knowledge of these chemistry principles will help you to better understand the benefits and hazards to mankind and enable you to make intelligent decisions in the future. The door 'Chemistry Grade 10' is open to you. In your career it will not be in vain, it will become a beneficial asset.

EXERCISES

1. Match each of the items given in List A with the appropriate correct item shown in List B.

List A	List B
(a) Aristotle	(i) discovered the gas, oxygen
(b) Robert Boyle	(ii) postulated the atomic theory
(c) Joseph Priestley	(iii) laid the background for a more quantitative approach to chemistry
(d) Antoine Lavoisier	(iv) very early chemist tried to turn cheaper metals to gold
(e) John Dalton	(v) explained the Law of Conservation of Matter
(f) Amedeo Avogadro	(vi) formulated the theories on the behaviour of matter
(g) Alchemist	(vii) developed the basic ideas about the behaviour of gases

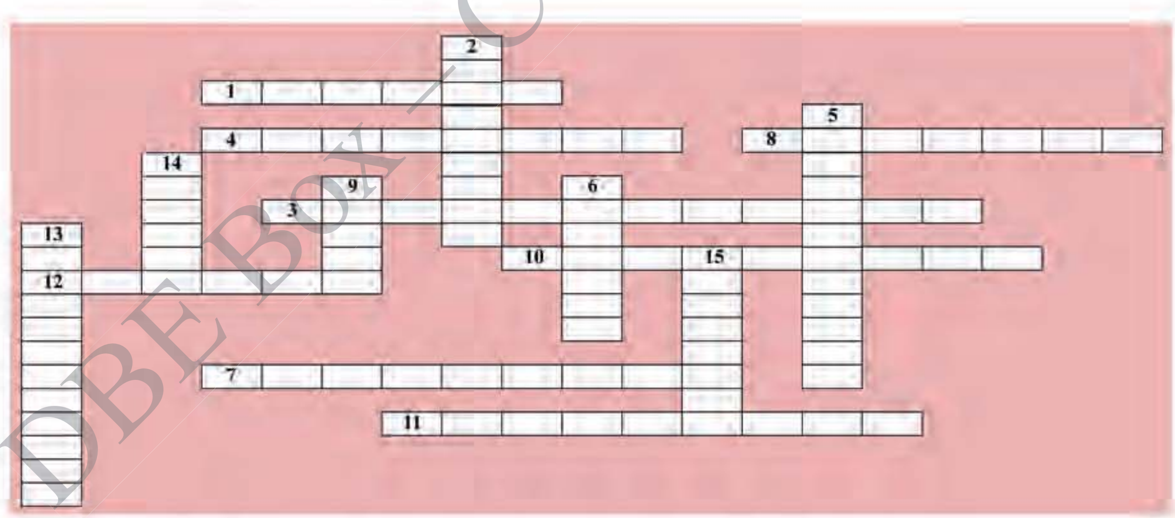
2. Fill in the blanks with suitable words and apply these words to solve the puzzle given below.

Across

(1) Chemistry is a study of _____ and their changes.
(3) Organic chemistry is an area of study on _____ and their derivatives.
(4) The behaviour and changes of matter and the related energy are studied in _____ chemistry.
(7) An area of study on matter not deal with hydrocarbons is _____ chemistry.
(8) Plastic is a synthetic _____ used to make a variety of products such as water plastic bottles.
(10) Chemical _____ occur when you breathe, eat, sit or read.
(11) Paper is made from _____.
(12) The air you breathe is a mixture of _____, nitrogen, etc.

Down

(2) Every activity, i.e., breathing, eating, reading, sitting, involves _____.
(5) Analytical chemistry is a study of components and _____ of substances.
(6) Chemistry is also considered as an experimental _____ since it is based on the results of observations and analyses through many investigations.
(9) A synthetic polymer, ____ is used for clothing.
(13) A study of matter and processes of living organisms is known as _____.
(14) Burning gasoline gives _____ to propel automobiles.
(15) Fermenting wine is _____ change.

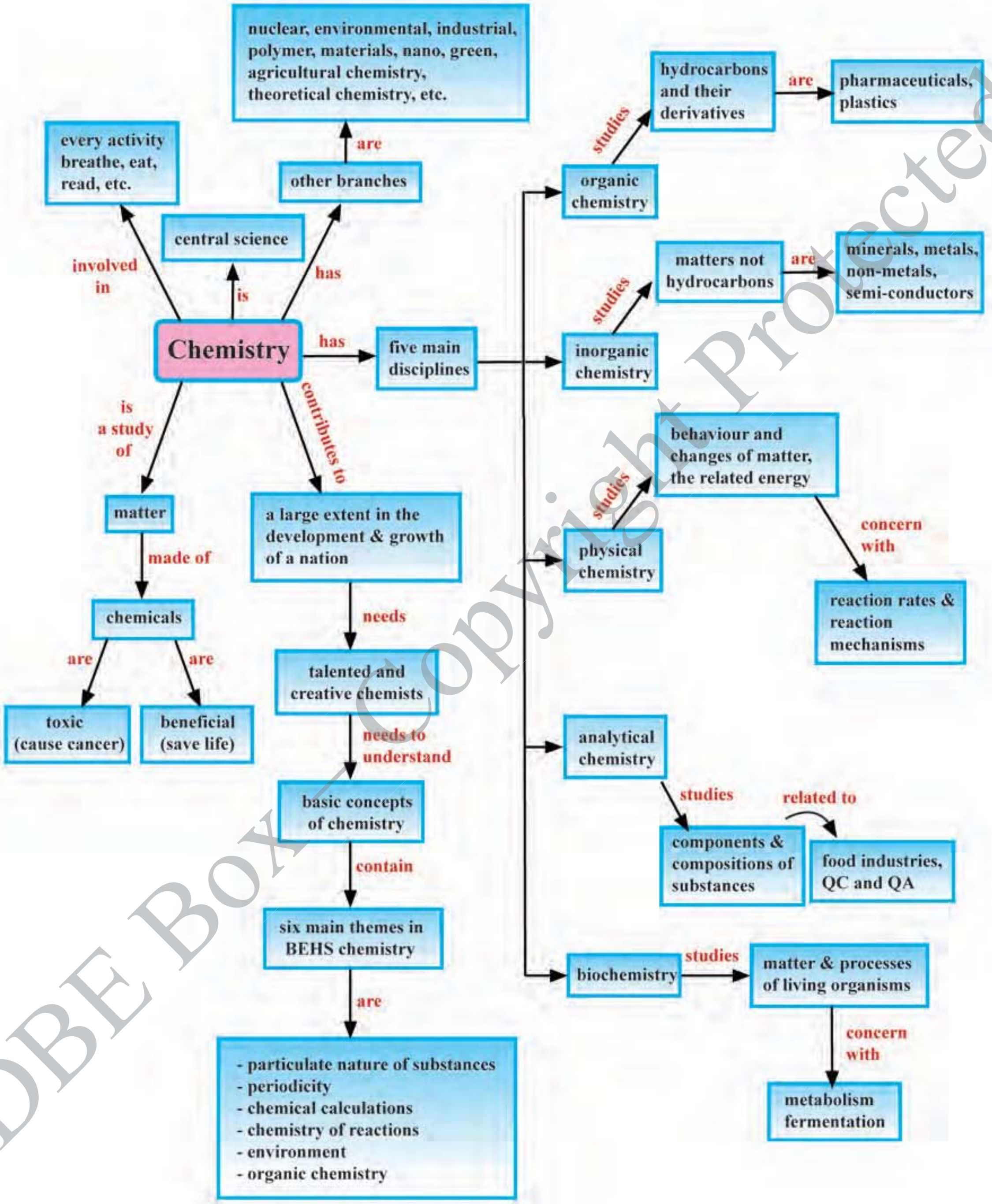


3. How does the study of chemistry relate to other areas of study in science?

4. In what ways does chemistry affect your life?

CHAPTER REVIEW

(Concept Map)



CHAPTER 2

MATTER AND SOLUTIONS

Chemistry is after all the study of all matter, its composition, its properties, and its transformation from one form to another. **What is matter?** This word is used to cover all the substances and materials from which the physical universe is composed. There are many millions of different substances known and all of them can be categorised as solids, liquids or gases. Nowadays, there is a fourth state known as 'plasma' which is a hot ionised gas containing charged particles.

Learning Outcomes

After completing this chapter, students will be able to:

- explain the theory of matter;
- identify the states of matter based on the arrangement and movement of atoms and molecules;
- analyse the changes of state based on changes in the arrangement and movement of atoms and molecules and the level of energy;
- discuss the characteristics of and the distinctions between elements, compounds and mixtures;
- differentiate between physical and chemical changes;
- compare the characteristics of, behaviours of and connections between a solute, solvent and solution;
- solve the solubility of substances and the effect of temperature on it;
- describe separation techniques as applied to mixtures.

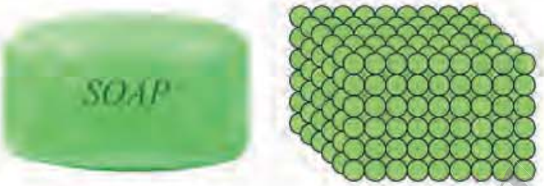
All substances are **matter**. Matter is made up of tiny particles. These can be atoms or molecules (groups of atoms), and elements or compounds. This includes the air, the sea, the Earth, all living creatures and even the galaxies.

The air (gas), the sea water (liquid), and alloys (solid) are not pure substances; they are mixtures. Many mixtures contain useful substance mixed with unwanted material. In order to obtain these useful substances, chemists often have to separate them from impurities. Different methods of separation depend on whether the substances to be separated are solids, liquids or gases. Adding sugar to tea or coffee is a solid-liquid mixture. This type of process involves solute, solvent and solution. What other examples can you think of where this type of process takes place?

2.1 STATES OF MATTER AND ARRANGEMENT OF PARTICLES IN MATTER

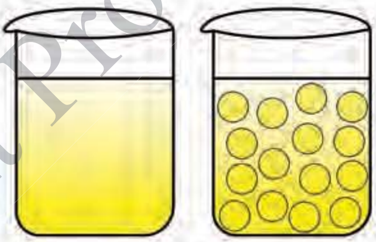
The most common states of matter are solid, liquid and gas. Water is a substance, which exists in all three states of matter: ice (solid), water (liquid) and steam (gas). The properties of each state of matter depend on the forces of attraction between the particles which can be weak or strong.

(a) Solids



The particles in a solid are packed very tightly together with strong forces between one another. Therefore, they have little freedom of movement and can vibrate about a fixed position. Solids have a definite shape and volume. They have different colours and different properties. Some solids are hard while others are soft. Some are dense while others are light. However, all solids have common properties, i.e., unlike gases they cannot be compressed and do not flow. Solids do expand slightly when heated.

(b) Liquids

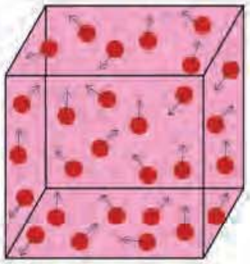


Liquids are composed of particles which are not fixed in any definite positions as in the solids. The particles are able to move freely throughout the liquids but not as independently as the gases so that the liquid can take up the shape of the container. Liquids can have a definite volume, because the particles in liquid are held together more strongly than those in gases. Some liquids have colours. However, unlike gases they cannot be compressed but they can flow easily.

(c) Gases

The particles in a gas are in constant and rapid motion because of weak attractive forces between gaseous particles. They move freely in all directions until they hit the walls of the container. The gas in a container spreads out to occupy the whole space of the container taking its shape and volume. Hence, gases do not have a definite volume and shape of their own.

Gases can have different properties. Some gases have a smell while others are odourless. Some gases have colours while others are colourless. However, all gases can easily be compressed and spread in all directions. It can be clearly visible in particles of smoke suspended in a gas. It is because of the effect known as **Brownian motion.**



(d) Particles of Matter: Diffusion Process

Diffusion can be explained by the Brownian motion.

The particles of matter are too small to be seen directly. It can be explained by indirect ways to show that matter consists of particles. One method is by a process called diffusion.

In diffusion, the particles of one substance mix with and move through the particles of another substance. Diffusion is mainly seen in liquids and gases. Some examples are given below:

A drop of food colouring matter diffuses throughout the water in a glass so that, eventually, the entire content of the glass will be coloured.

When a few drops of perfume are released into a room, the particles of perfume move through the air and spread the room. Anyone in the room would be able to smell the perfume eventually.

The speed of diffusion of particles is affected by the mass of particles and by the temperature.

- The bigger the mass, the slower the particles diffuse. The smaller the mass, the faster the particles diffuse.
- The higher the temperature, it enables the particles to diffuse faster.

Chemistry in Daily Life

- Matter is everything that we come across in our lives, like the air you breathe, the clothes you wear and the water you drink.
- The most common states of matter are solid, liquid and gas. Some are in solid (ice, sugar, salt, iron, copper, etc.), some are in liquid (water, oil, juice, etc.), some are in gaseous (air, oxygen, carbon dioxide, etc.) states.
- Diffusion of particles can occur in many ways. Preparing tea using tea bag in hot water, smelling perfume, aroma of foods, aroma therapies etc. are some examples of diffusion.

Review Questions

(1) Distinguish among the solid, liquid and gas.

(2) Which states can you see the following matter in our environment as solid or liquid or gas?
(a) iron (b) water (c) mercury (d) argon (e) gold (f) copper (g) vinegar

Key Terms

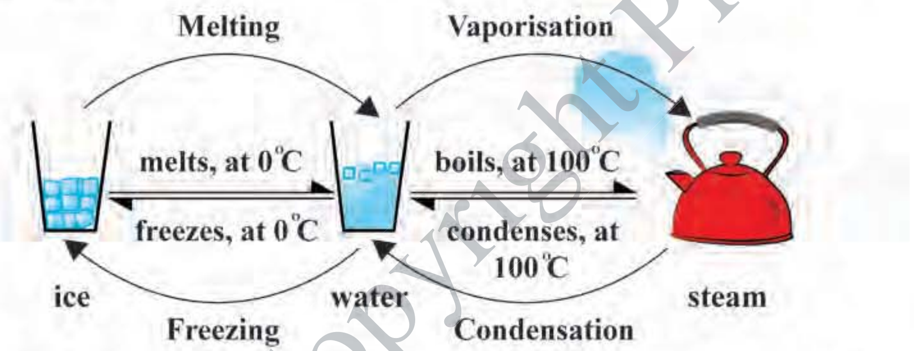
- **Matter** is made up of tiny particles, and has mass and takes up space. Three common states of matter are solid, liquid and gas.
- **Brownian motion** is the continuous random movement of small particles suspended in a gas or liquid, which arises from collisions with the gas or liquid particles, e.g., the motion of pollen grains on still water, movement of invisible dust in a room.

2.2 CHANGES IN MATTER

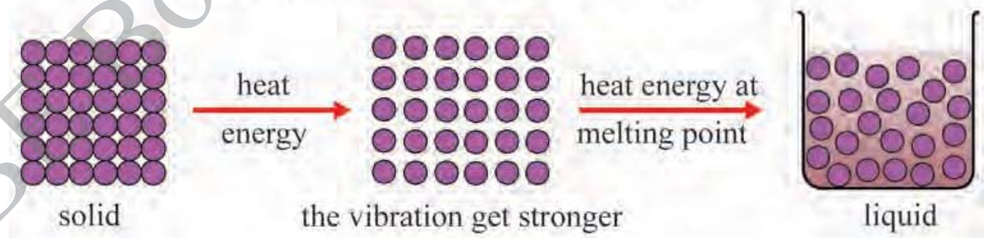
The materials around us are subject to constant change. Plants and animal materials decay, metals corrode, and land areas erode. Moreover, every substance – for example, water, sugar, salt, gold or silver – has a set characteristics or properties that distinguish it from all other substances and gives it a unique identity. One way to classify properties is based on whether or not chemical composition of an object is changed by the act of observing the property. Changes in substances can be classified as either physical or chemical.

(a) Physical Changes

A physical change is a change in which no new substances are formed. For example, when ice melts from solid to liquid, or when sand is ground to a fine powder, no new substance is formed. Melting, boiling, freezing, evaporation, vaporisation, condensation and sublimation are considered as physical changes.



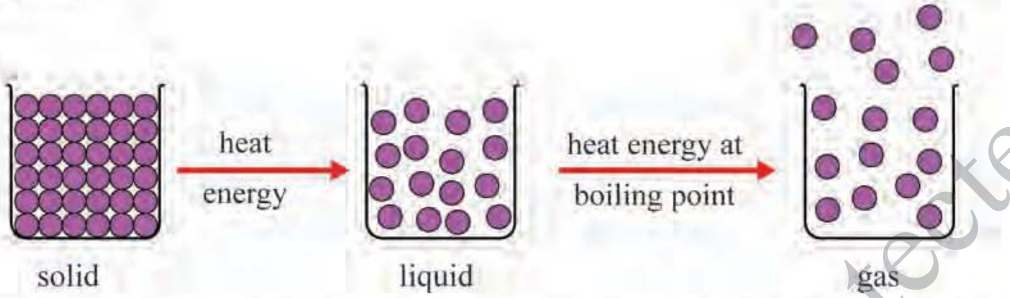
In a solid the particles attract one another. There are attractive forces between the particles which hold them close together. The particles have little freedom of movement and can only vibrate about a fixed position. When a solid substance is heated, the particles of the solid vibrate more strongly. Finally, these particles are able to overcome the forces that hold the particles in their fixed position. The solid then turns into a liquid. This is called melting. The temperature at which a solid melts is called the melting point.



When heating is continued, the particles of the liquid gain more energy and move more quickly as the temperature is increased. Eventually, the particles have enough energy to completely break the forces holding them together. Then the liquid particles escape from the surface to the space above the liquid. The particles are now able to move freely and get farther. A gas is formed. This is called vaporisation. **Vaporisation** is the process that occurs when a chemical or element is converted from a liquid to a vapour.

Evaporation is the process of a substance in a liquid state changing to a gaseous state due to an increase in temperature and / or pressure.

Boiling is the rapid vaporisation of a liquid, which occurs when a liquid is heated to its boiling point.



An unusual change of state can also occur, i.e., change of solid state to gaseous state and vice versa. For example, iodine solid changes to iodine vapour, where heat is absorbed. This is known as **sublimation**. Water vapour turns to frost is an example of **deposition**.

You can readily see that there is no new chemical substance formed in these changes. It is known as physical change; for example, sugar being dissolved in water, moulding silver and gold, passing electricity in electric bulb.

(b) Chemical Changes

A chemical change is a change in which one or more new substances are formed. Examples of chemical changes are cooking of rice from rice grains, green mangoes ripening, burning of a match, and burning of candle.

Chemical changes occur via chemical reactions such as dissociation, neutralisation, precipitation, etc. For example,

- Heating of limestone or marble (dissociation or decomposition)
- Use of magnesia to treat gastric patient (neutralisation)
- Passing carbon dioxide into limewater (precipitation)

Chemistry in Daily Life

- Freezing, melting, boiling and dissolving, cutting and moulding of the substances, etc. are physical changes.
- Mothballs sublime into vapour of deodorant for toilets and bathrooms. It is also a physical change.
- In our daily life, cooking the foods, burning the candle, iron rusting, vegetables rotting, building a fire, photosynthesis reaction, making soaps and detergents, etc. are chemical changes.

Review Questions

(1) Is squeezing juice from lime a physical change or chemical change?
Give reason for your answer.

(2) Classify the following changes into physical change and chemical change:
  (a) boiling an egg
  (b) mixing sand and water
  (c) making jelly
  (d) evaporating alcohol
  (e) souring of milk
  (f) baking a cake
  (g) digesting food
  (h) crushing a can
  (i) breaking a glass
  (j) mixing green and red marbles

(3) State the change of each of the following processes:
  (a) When iodine, a solid, is gently heated it forms directly into a purple gas.
  (b) Frost is formed when water vapour is cooled.
  (c) Lime water becomes cloudy when carbon dioxide gas is passed into it.

Key Terms

- The temperature at which a solid changes to a liquid state at one atmospheric pressure is called **melting point** of that solid.
- The temperature at which the vapour pressure of the liquid is equal to the atmospheric pressure of the surrounding is called the **boiling point**.
- **Vaporisation** is the process that occurs when a chemical or element is converted from a liquid to a vapour.
- **Evaporation** is the process of a substance in a liquid state changing to a gaseous state due to an increase in temperature and / or pressure.
- **Freezing** is the process in which a liquid becomes sufficiently cold to change into a solid. **Freezing point** is the temperature at which a liquid becomes a solid.
- **Condensation** is the change from a gaseous state to its liquid state.
- **Sublimation** is the change of solid state directly into gaseous state without melting.
- **Deposition** is the direct solidification of a vapour by cooling; the reverse of sublimation.
- A **physical change** is a change in which no new substances are formed. There may be a temporary change in colour, temperature and state of the substances but no new substances are formed in the physical change.
- A **chemical change** is a change in which one or more new substances are formed. The substances change in colour, temperature and state but they also change into a new substance or substances in the chemical change.

2.3 ELEMENTS, COMPOUNDS AND MIXTURES

All samples of matter can be divided into two categories: pure substances and mixtures. A pure substance is a form of matter that always has a definite and constant composition. Pure substances are classified as either elements or compounds. At the beginning of the 19th century, John Dalton proposed the theory of matter: that all matter was composed of atoms, which were invisible and indivisible. Today, the atom is still considered as the basic unit of any element. An atom may combine chemically to form molecules; the molecules become the smallest unit of any substances that possesses the properties of that substance. Modern experimental evidence has shown that atoms are divisible to create either lighter or heavier atoms.

(a) Elements

Elements are substances consisting of one type of atom, e.g., carbon element is made up of carbon atoms. Atoms are the smallest particles into which an element can be divided.

There are 92 known elements which occur naturally, either in the free or combined state. Some elements are solids such as copper, iron, zinc, silver, gold, carbon and phosphorus. Some elements are liquids. They are mercury and bromine. Some elements are gases such as oxygen, nitrogen, hydrogen and chlorine. Substances like these, which cannot be broken down into a simpler substance by chemical means, are called elements.

On the basis of their properties, elements may be classified into two groups, metals and non-metals (Table 2.1).

Table 2.1 General Properties of Metals and Non-metals

Metals	Non-metals
Metals show metallic luster.	Non-metals do not show metallic luster.
Metals have high density.	Non-metals have low density.
Most of the metals are malleable and ductile.	Non-metals are usually brittle.
Metals are good conductors of heat and electricity.	Most of the non-metals are poor conductors of heat and electricity.

(b) Compounds

The atoms of some elements are joined together in small groups. These small groups of atoms are called molecules.

Molecules exist in elements as well as compounds. A molecule of an element (molecular element) consists of atoms of the same kind. A molecule of a compound (molecular compound) consists more than one kind of atoms. The atoms of different elements in the molecule of a compound are combined in a definite ratio.

Most substances on Earth occur as compounds, e.g. carbon dioxide (CO_2), water (H_2O), marble ($CaCO_3$), glucose ($C_6H_{12}O_6$), ethanol (C_2H_5OH) and ammonia (NH_3). Although there is only small number of elements, there are millions of compounds.

Formulae and types of some compounds are described in Table 2.2.

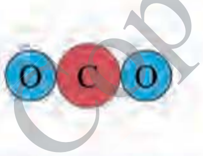
Two or more different elements may combine together to form compounds. Some compounds occur naturally but some are made in laboratories, e.g., water occurs in nature and ethanol is a man-made compound.

The compounds can be classified in various ways. They can also be classified based on the combination of the number of atoms or the number of different elements. Hydrogen molecule is formed by two atoms of hydrogen. So, it is a diatomic molecule. Water is formed by combining two different elements: H and O. So, it is a triatomic molecule (binary compound).

Carbon dioxide is formed by the combination of two different elements, C and O. So it is a triatomic molecule (binary compound). In ammonia, there are more than three atoms from two different elements: N and H. So, it is a polyatomic molecule (binary compound). In ethanol, there are more than three atoms from three different elements: C, H, O. Therefore, it is a polyatomic molecule (ternary compound) (Table 2.2).

The most obvious difference is that an element cannot be broken down into other substances by chemical means whereas a compound can be broken down into other substances by chemical means.

Table 2.2 Name, Formula and Type of Compounds

Name	Element	How the atoms are joined	Formula	Type of compound (based on number of atoms)	Occurrence
water	hydrogen and oxygen		H_2O	triatomic molecule (binary compound)	natural
carbon dioxide	carbon and oxygen		CO_2	triatomic molecule (binary compound)	natural
ammonia	nitrogen and hydrogen		NH_3	polyatomic molecule (binary compound)	man-made
ethanol	carbon, hydrogen and oxygen		C_2H_5OH	polyatomic molecule (ternary compound)	man-made

(c) Mixtures

Mixtures consist of two or more different substances that are mixed physically but not chemically combined. They do not have well defined specific properties and the substances are not in fixed ratios.

The substances in a mixture may be solids, liquids or gases. For example, brass, a solid, is a mixture of the elements copper and zinc; sea water is a mixture of compounds including mainly water and sodium chloride; air is a mixture of gases containing nitrogen, oxygen, argon, carbon dioxide and water vapour. The mixtures may also be heterogeneous or homogeneous (Table 2.3). Therefore, the mixtures can be classified as two main categories: homogeneous and heterogeneous mixtures.

Table 2.3 Different Types of Mixtures

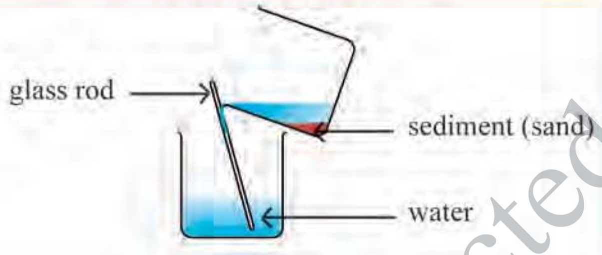
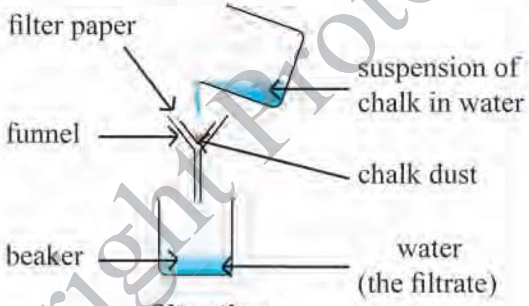
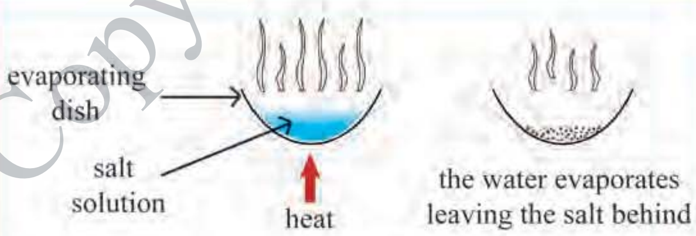
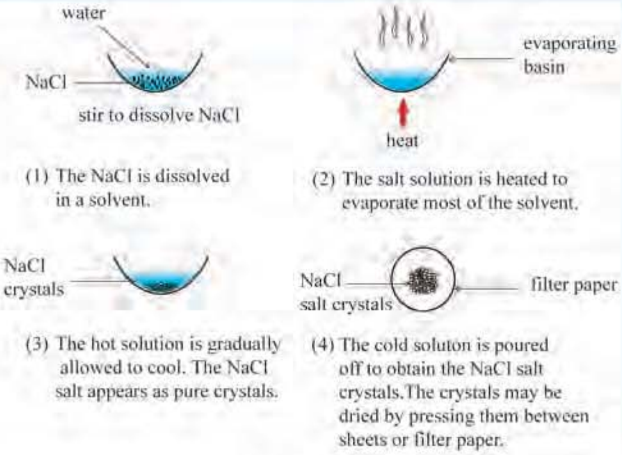
Physical state	Type of mixture	Example
solid-solid	homogeneous	stainless steel (mixture of iron and chromium)
solid-solid	heterogeneous	flour and rice powder
solid-liquid	homogeneous	sugar solution (sugar and water)
solid-liquid	heterogeneous	salt and oil
solid-gas	heterogeneous	dust in air
liquid-liquid	homogeneous	vinegar (mixture of acetic acid and water)
liquid-liquid	heterogeneous	oil and water
liquid-gas	homogeneous	soft drink (carbon dioxide gas dissolved in sterilised water at high pressure)
liquid-gas	heterogeneous	fossil fuel (mixture of crude oil and natural gas)
gas-gas	homogeneous	air (mixture of different gases)

Apart from **alloys** containing two metals, solid-solid mixtures are heterogeneous. Some heterogeneous mixtures cannot be recognised by the naked eyes, such as a mixture of magnesium oxide and calcium oxide. However, solid-solid mixtures can be recognised by microscopic examination, whereas we cannot do so with homogeneous solutions. Therefore, we are able to differentiate between a homogeneous mixture and a heterogeneous mixture by visual examination.

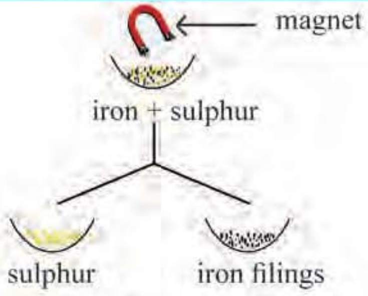
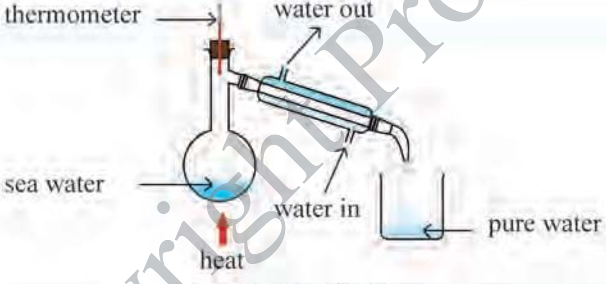
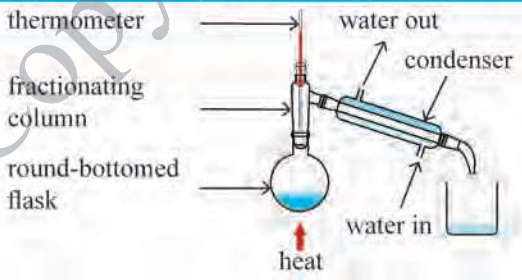
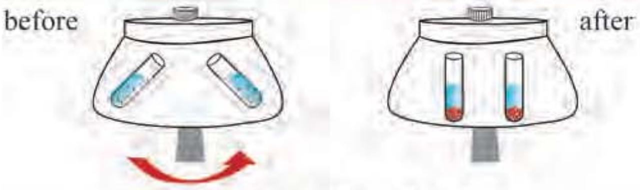
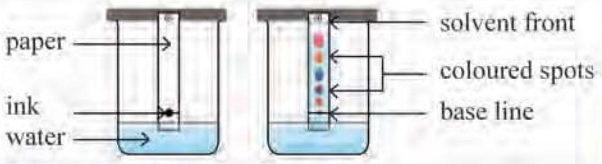
(d) Separation of Mixtures

Most substances are naturally found as mixtures; therefore the separation methods shown in Table 2.4 indicate how the physical states of components in the mixture can be separated into pure substances.

Table 2.4 Some Separation Methods of Mixtures

Types of mixtures	Separation method
a liquid and a solid mixture such as a suspension (sand/water)	 **decantation**
a solid from a liquid (chalk dust from water)	 **filtration**
solute from its solution (sodium chloride, NaCl salt from its solution)	 **evaporation**
a solute crystal from its solution (sodium chloride, NaCl salt from its solution)	 (1) The NaCl is dissolved in a solvent. (2) The salt solution is heated to evaporate most of the solvent. (3) The hot solution is gradually allowed to cool. The NaCl salt appears as pure crystals. (4) The cold soluton is poured off to obtain the NaCl salt crystals.The crystals may be dried by pressing them between sheets or filter paper. **crystallisation**

Continued from **Table 2.4**

Types of mixtures	**Separation method**
to attract magnetically susceptible materials (sulphur and iron mixture)	 **magnetic separation**
a solvent from a solution (pure water from sea water)	 **simple distillation**
liquids from each other (separation of petroleum)	 **fractional distillation**
a solid from a liquid (milk, blood)	 **centrifugation**
different substances from a solution (separation of ink by paper chromatography)	 **chromatography**

Chemistry in Daily Life

- Gold, silver and copper have been used to make ornamental objects and jewellery for thousands of years. Special properties of gold make it perfect for manufacturing jewellery.
- Mercury is used in thermometers and traditional blood pressure monitors. Mercury is a very toxic substance. When it accumulates in the body, it causes damage to the brain, kidney and lungs.
- Calcium carbonate is a compound made up of three elements-calcium, carbon and oxygen. Chalk is one form of calcium carbonate.
- Examples of homogeneous mixtures include sugar solution, which is the mixture of sucrose and water, and gasoline which is a mixture of dozens of hydrocarbon compounds.
- Colloidal mixtures (heterogeneous) have components that tend not to settle out. Milk is a colloid of fat globules suspended in water.
- The mixture of gasoline / kerosene and water is an example often cited as a safety hazard. Decanting a mixture containing flammable solvents can be dangerous as the flammable material evaporates and forms dangerous fumes.
- When rain touches the ground it mixes with dirt, rocks and so on, i.e., it could become a mixture.
- Sediment from the fermentation process of wine can produce an undesirable taste. Wine is separated from the sediments by decantation.
- Plasma can be removed from blood by decantation after centrifugation.

Review Questions

(1) When attempts are made to break down substance **A** by chemical methods, the same original substance is always formed. Is substance **A** an element or a compound?

(2) When a substance is broken down by chemical means, two substances with different properties are formed. Is the original substance an element or a compound?

Key Terms

- An **element** is a substance that cannot be broken down into other simpler substances through chemical means. Every element is made up of its own type of atoms. Therefore, it has a unique position in the Periodic Table.
- A **compound** is a substance containing two or more different elements chemically joined together in a fixed ratio.
- A **molecule** is the simplest unit of the chemical substance, usually a group of two or more atoms.
- **Molecules** exist in elements as well as compounds. A molecule of an element (molecular element) consists of atoms of the same kind. A molecule of a compound (molecular compound) consists of more than one kind of atoms. The atoms of different elements in the molecule of a compound are combined in a definite ratio.

- **Diatomic molecules** are molecules composed of only two atoms of same or different elements.
- **Triatomic molecules** are molecules composed of only three atoms of same or different elements.
- **Polyatomic molecules** are molecules composed of three or more atoms of same or different elements.
- The compounds formed by the combination of two elements are called **binary compounds**.
- The compounds formed by the combination of three elements are called **ternary compounds**.
- A **mixture** is a combination of more than one substance, where these substances are not bonded to each other. It consists of two or more substances which may be present in any proportion by weight. The constituents of the mixture do not combine chemically.
- A **heterogeneous mixture** is one that is non-uniform, and where the different components of the mixture can be seen. The components separate, and the composition varies.
- A **homogeneous mixture** is one in which the composition of its components are uniformly mixed throughout. The components cannot be seen separately on visual or microscopic examination.
- **Alloy** is a substance made by combining two or more metallic elements, especially to give greater strength or resistance to corrosion.
- **Filtration** is a method for separating an insoluble solid from a liquid. When a mixture of sand and water is filtered, the sand remains as residue on the filter paper and the water, which is also called filtrate, passes through the filter paper.
- **Crystallisation** is defined as a process by which a chemical is converted from a liquid solution into a solid crystalline state.
- **Decantation** is a process to separate mixtures. Decanting is just allowing a mixture of solid and liquid or two immiscible liquids to settle and separate by gravity.
- **Magnetic separation** is used to separate the components of a mixture when at least one of them is magnetic in nature.
- **Simple distillation** is a procedure by which two liquids with different boiling points can be separated. It is used to separate solvent from a solution.
- **Fractional distillation** is a method for separation of a liquid mixture into fractions with different boiling points (and hence chemical composition) by means of distillation, typically using a fractionating column.
- **Centrifugation** is a technique used for the separation of particles from a solution according to their size, shape, density, viscosity of the medium and rotor speed.
- **Chromatography** is a separation method of the mixed substances that depends on the speed at which they move through special media, or chemical substances. It consists of a stationary phase (a solid) and a mobile phase (a liquid or a gas).

2.4 SOLUTIONS AND SOLUBILITY

A solution is a homogeneous (uniform) mixture of two or more substances. The ocean is a vast water solution containing different compounds extracted from the minerals of the Earth's crust. Nutrients are carried in water solution to all parts of a plant.The production of many useful materials by the chemical industry involves chemical reactions in which the reacting substances are dissolved in water, ethanol (ethyl alcohol), etc. Solutions play an important part in many processes that go on about us.

(a) Solutions

Some solids such as copper(II) sulphate, sugar and common salt are soluble in water but some solids such as sand, charcoal and chalk are insoluble in water. Some solids such as iodine are slightly soluble in water. When you mix sugar with water, it seems to disappear. That is because its particles spread all through the water particles. The sugar has dissolved in the water, giving the mixture called as a **solution**. Sugar is the **solute**, and water is the **solvent**.

Some liquids can mix with one another in all proportions (miscible liquids), while some liquids do not mix (immiscible liquids). Thus for example, ethanol, acetic acid and sulphuric acid are soluble in water but petrol and oils are insoluble in water.

Some gases such as hydrogen chloride and ammonia are very soluble in water. Some gases such as oxygen, nitrogen and hydrogen are not very soluble in water.

Solutions may be gaseous, solid, or liquid in nature. Dry air is a familiar example of a gaseous solution. Brass (copper and zinc) is an example of a solid solution. Liquid solutions may contain solid, liquid, or gaseous solutes. Salt water is a familiar example of a solid dissolved within a liquid. Vinegar is a solution containing two liquids, acetic acid and water. Carbonated water contains carbon dioxide gas molecules existing between molecules of water.

Solutions having water as the solvent are referred to as aqueous solutions. Many reactions including those vital for life processes occur in aqueous solutions. Blood and saliva are some of the more familiar solutions of biological importance.

(b) Solubility

Some solid is added to a certain volume of water in a beaker with stirring until all of it dissolved. If some more solid added can dissolve with stirring, this solution is known as an **unsaturated solution**. A **saturated** solution forms when no more solute can be dissolved in the given amount of solvent at that temperature even if it is stirred. The amount (in grams) of the solute in 100 g of water to give a saturated solution at that experimental temperature is known as the **solubility** of that solute. If the solution contains more solute than it should have at room temperature, it is a **supersaturated** solution.

Effect of temperature on solubility

Solubility of the solute depends, in part, on the temperature of the solvent. When the temperature of the solution is increased the solubility of the solute increases (Table 2.5 and Figure 2.1).

Table 2.5 The Solubility of Potassium Nitrate in Water at Different Temperatures

Temperature (°C)	10	20	30	40	50	60	70	80	90	100
Solubility (g / 100 g water)	21.2	31.6	45.3	61.4	83.5	106	135	167	203	245

From the solubility data of the potassium nitrate in the above table, a graph is drawn with temperature on the horizontal axis and solubility on the vertical axis as shown in Figure 2.1.

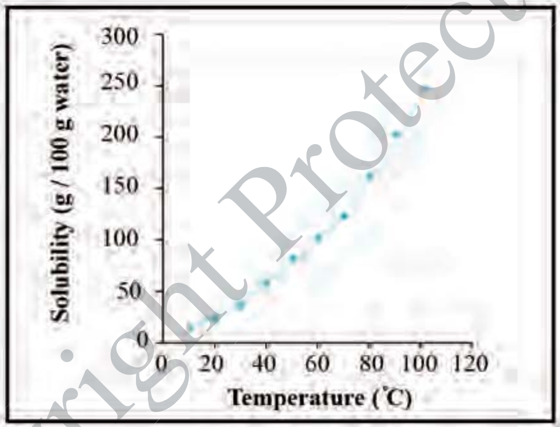


Figure 2.1 The Solubility of Potassium Nitrate in Water at Different Temperatures

There are some exceptions. The solubility of all gases and some solids such as calcium hydroxide and calcium sulphate decreases when temperature of the solution is increased. It is, therefore, necessary to specify the temperature at which the solution is saturated. A good example is opening two cans of soda, one cold one and one warm one. Comparing the reactions will demonstrate that more gas is released from the warm pop than from the cold pop.

When a saturated solution at a higher temperature is cooled, the solubility decreases. So the excess solute, i.e., the difference in the solubility of the two temperatures, will come out as solid.

Chemistry in Daily Life

- Solution can be found almost everywhere on the earth, from the oceans to the sky. Every ocean and every lake on earth is a solution, because the water has mixed with dirt, salt and various substances.
- When you stir sugar in a cup of coffee, you are making a solution.
- Solubility has many practical applications in our lives such as purifying water and making drinks.

Review Questions

(1) Give two examples for each of the following:
  (a) solids that dissolve in water (b) insoluble solids in water
  (c) solvents other than water

(2) 20 g of a soluble substance is dissolved in water to form 100 g of the solution. 25 g of the solution is taken and evaporated to dryness. How many grams of the solid will be obtained?

(3) The solubility of copper(II) sulphate at 60 °C is 40 g / 100 g, and at 90 °C is 67.5 g / 100 g.

A saturated solution of copper(II) sulphate in 100 g of water at 90 °C is cooled to 60 °C. Calculate the amount of copper(II) sulphate which would come out of the solution.

Key Terms

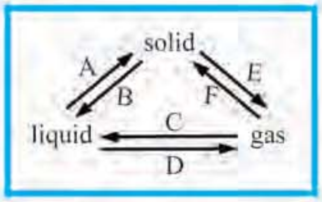
- A **solute** is a substance which dissolves in a solvent to give a solution.
- A **solvent** is a substance, mostly liquid, in which another substance dissolves to give a homogeneous mixture.
- A **solution** is a clear homogeneous mixture obtained when a substance dissolves in a solvent. In a solution the solute is uniformly distributed throughout the solution.
- **Solubility** of a substance at a given temperature is the mass in grams of the substance which will saturate 100 g of water, at that temperature.
- A **saturated solution** is one in which no more solute will dissolve at the given temperature, in the presence of excess solute. A solution in which more of the solute can dissolve at the given temperature is called **unsaturated solution**.
- The solution that retains more solute than that required to saturate the solution at room temperature is called a **supersaturated solution**.

EXERCISES

1. This question is about ways to separate and purify substances. Match each term from List A with the correct description from List B.

List A	List B
(a) evaporation	(i) a solid appears as the solution cools
(b) condensing	(ii) used to separate a mixture of two liquids
(c) filtering	(iii) the solvent is removed as a gas
(d) crystallising	(iv) this method allows you to recycle a solvent
(e) distillation	(v) a gas changes to a liquid, on cooling
(f) fractional distillation	(vi) separates an insoluble substance from a liquid

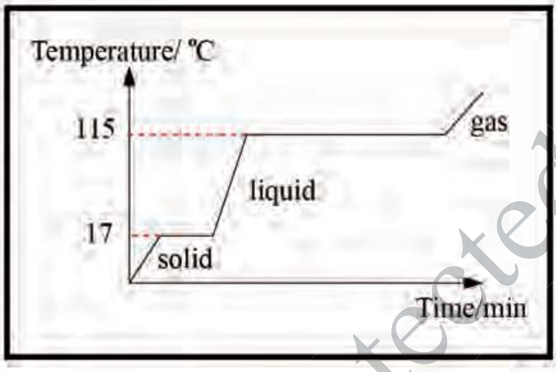
2. The following diagram shows the three states of matter and how they can be interchanged.
   (a) Name the changes of A to F.
   (b) Name a substance which will undergo changes from solid to liquid to gas between 0 °C and 100 °C.
   (c) Describe what happens to the particles of the solid during change E.
   (d) Name a substance which will undergo change E.



3. (a) Why solids do not undergo diffusion? Explain why diffusion of gases is faster than liquids.
   (b) Give two examples for diffusion of gases and liquids found at home.

4. When a jar of coffee is opened, people in all parts of the room soon notice the smell. Explain how this happens.

5. The heating curve for a pure substance is given. It shows how the temperature rises over time, when the substance is heated until it melts, then boils.
   (a) What is the melting point of the substance?
   (b) What happens to the temperature while the substance changes state?
   (c) The graph shows that the substance takes longer to boil than to melt. Suggest a reason for this.
   (d) How can you tell that the substance is not water?
   (e) Sketch a rough heating curve for pure water.



6. The solubility (g / 100 g water) of three substances at different temperatures are given below.

Temperature °C	0	10	20	30	40	50	60	70	80
Potassium chlorate	3.3	5.0	7.3	10.0	14.0	18.5	24.0	30.2	37.5
Calcium hydroxide	0.13	0.13	0.12	0.11	0.10	0.00	0.08	0.06	-
Sodium sulphate	5.0	9.0	20.5	41.0	48.0	47.0	45.0	44.0	43.0

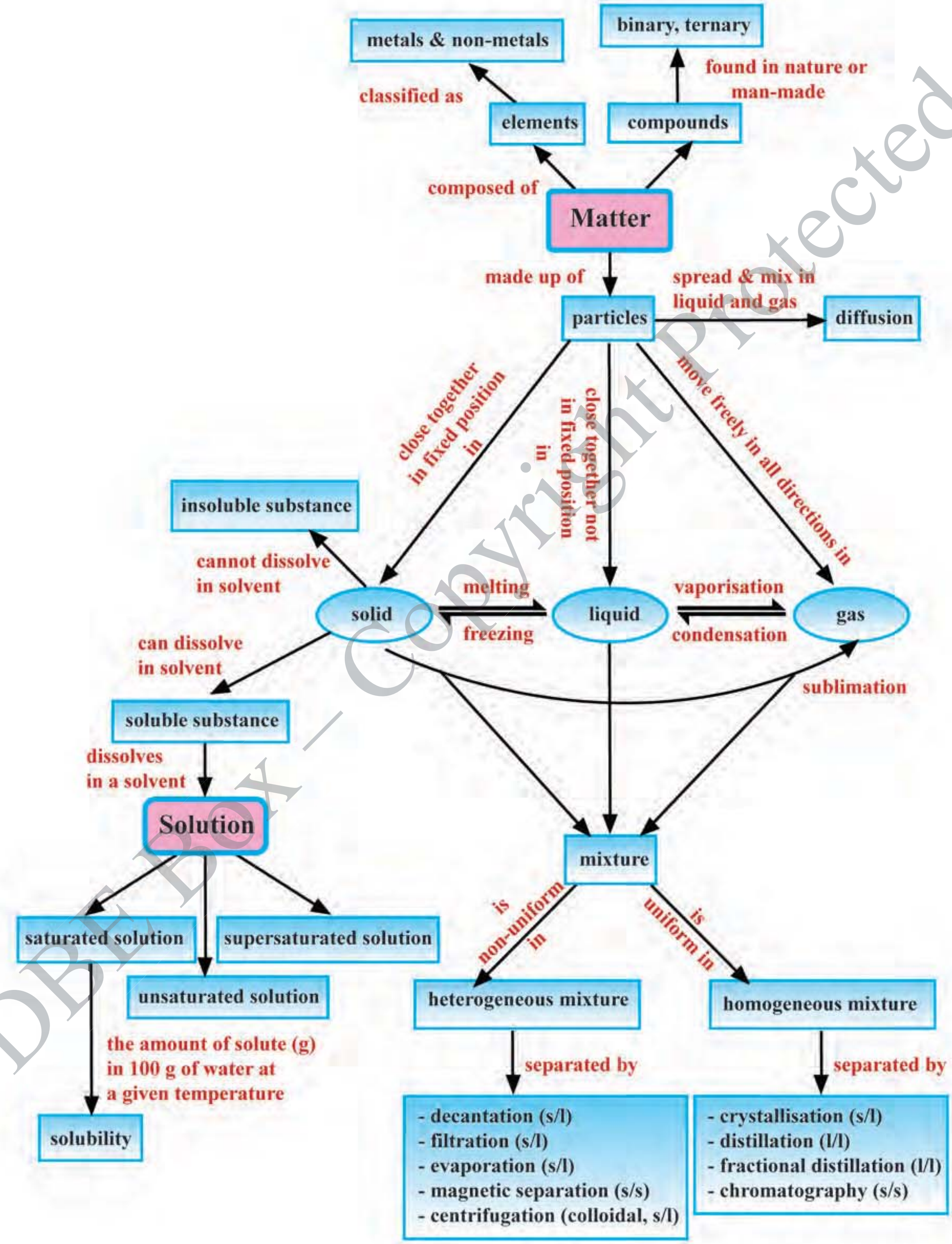
   (a) Plot the solubility curve of each substance.
   (b) Describe the change in solubility with the temperature for each substance.
   (c) What is the solubility of each substance at 25 °C?
   (d) What happens when each solution at 70 °C is cooled down to 30 °C?

7. The solubility of sodium nitrate at 40 °C is 104 g / 100 g water.
   (a) How much sodium nitrate will be obtained if 25.5 g of saturated solution at 40 °C is evaporated to dryness?
   (b) What is the maximum amount of solid that can be dissolved in 250 g of water at 40 °C?

8. The solubility of solid **A** at 60 °C is 24 g / 100 g water.
   (a) What is the amount of solid required to saturate 30 g of water at 60 °C?
   (b) What will be the amount of saturated solution obtained at 60 °C when 12 g of the solid **A** is used to prepare a saturated solution?

CHAPTER REVIEW

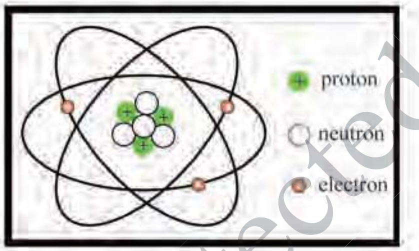
(Concept Map)

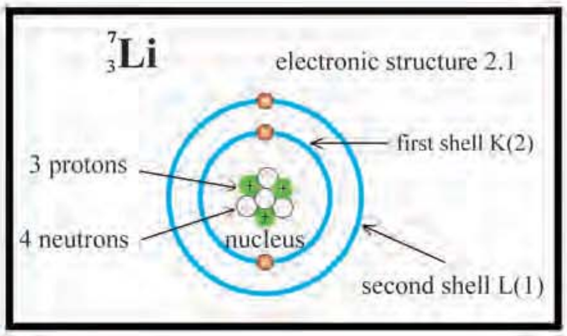


CHAPTER
3

THE ELECTRONIC STRUCTURES OF ATOMS AND PERIODIC TABLE

In Chapter 2, we have seen the word 'atom', which is the smallest particle of an element. In this chapter we will learn that the arrangement of atoms in the element can be explained by the new model. Atoms very rarely exist by themselves. They are usually joined together in groups by chemical bonds. Two types of chemical bonding: ionic and covalent, will be explained in this chapter.





We know that chemical elements play a vital role in our daily lives and are crucial for humankind, our planet, and our industries. The development of the Periodic Table of the elements is one of the most significant achievements in science. The Periodic Table gives insights into the elements and helps us to understand the characteristics of elements.

Learning Outcomes

After completing this chapter, students will be able to:

- describe the properties of electrons, protons and neutrons;
- define isotopes and isobars, and explain them with respect to atomic number and mass number;
- explain the electronic structure of atoms;
- explain how the Periodic Table is organised based on atomic structure;
- classify elements based on electron configurations;
- describe the periodic properties of common elements;
- discuss the general trends in metallic and non-metallic character, electronegativity, sizes, ionisation energy and electron affinity of elements within the periods and groups of the Periodic Table;
- compare the different types of bonds that form between atoms when molecules are formed.

3.1 STRUCTURE OF ATOM

All matter is made up of atoms. Atoms are so tiny that it was not realised that atoms were in fact made up of charged particles until about 1900 AD. The old model of the atom was published by Dalton in 1803. However, based on the experimental evidence, Dalton's atomic model has been replaced by the new model.

(a) Arrangement of Sub-atomic Particles in Atom

Atoms consist of three sub-atomic particles called proton, neutron and electron. They are also known as fundamental particles. The properties of fundamental particles of an atom are summarised in Table 3.1.

Table 3.1 Properties of Fundamental Particles

Particle	Symbol	Relative mass	Relative charge
proton	p	1	positive (+1)
neutron	n	1	neutral (0)
electron	e	1/1837	negative (–1)

A **proton** is a particle carrying a positive charge. A **neutron** is a particle carrying no charge and having a mass similar to proton. An **electron** is a particle carrying a negative charge and having a very small mass. The mass of an electron is approximately 1/1837 times the mass of a proton.

In a neutral atom, the number of protons is equal to the number of electrons. Atoms are identified by the number of protons in it. The **proton number** is referred to the number of protons in the atom and is often referred to as the **atomic number (Z)** of the element. Different elements have a different proton number or atomic number from each other.

The atom has a small, dense, positively charged center called **nucleus**. It contains proton(s) and neutron(s). The total number of these two particles is also known as the **nucleon number** or **mass number (A)**. Outside and around the nucleus, electrons rapidly move in circular or near circular orbits. It can be represented in Figure 3.1.

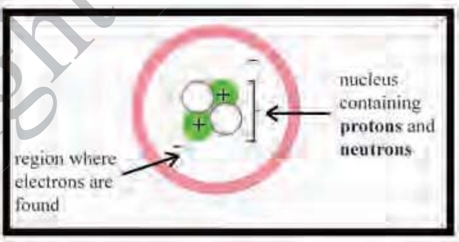


Figure 3.1 Atomic Structure of an Atom

We can describe any elements in a short way like this:

mass number A
X **symbol of element**
atomic number Z

For example, $^{12}_{6}C$, $^{16}_{8}O$ and $^{35}_{17}Cl$

(b) Isotopes

Not all atoms of an element are necessarily the same. All atoms of the same element have the same number of protons. However, some atoms of an element have different number of neutrons. For example, although all carbon atoms have six protons, not all carbon atoms are identical. Some have more neutrons than others.

Most of carbon atoms have 6 neutrons but some carbon atoms are found with 7 neutrons or 8 neutrons as shown in the diagram (Figure 3.2). These three different carbon atoms are called **isotopes** of carbon.

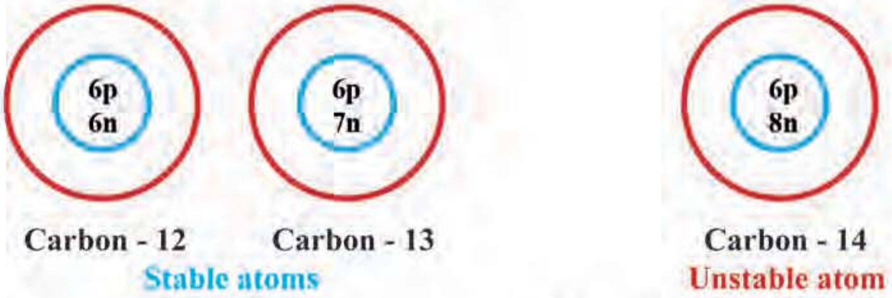


Figure 3.2 Isotopes of Carbon

Isotopes are atoms with the same number of protons but different number of neutrons. In other words, isotopes are the atoms of the same element with different masses.

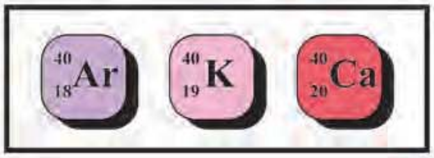
With the help of the mass spectrometer, some of the elements were found to consist of a mixture of the atoms having different masses. Mass spectrometer is a device used to find out the masses of isotopes and their relative abundance. It becomes a mean of measuring the ionisation energy of all elements.

Isotopes of the same element have the same chemical properties but there are some small differences in physical properties. For example, the three hydrogen isotopes, ${}^{1}_{1}H$, ${}^{2}_{1}H$ and ${}^{3}_{1}H$ have slightly different boiling points.

Some isotopes are **radioactive** whereas some are **non-radioactive**. A carbon-14 atom is radioactive but carbon-12 and carbon-13 are not radioactive. Radioactive isotopes produce radiation. Some radiations are harmful but some can be used for good. Most isotopes in the air and ground are not radioactive.

(c) Isobars

Some of the different elements have same nucleon number. That is, different elements with different proton number and same nucleon number. An example of a series of isobars would be ${}^{40}Ar$, ${}^{40}K$ and ${}^{40}Ca$.



Chemistry in Society

Up to the early 20th century, uranium was used in making an attractive yellow glaze for pottery. This was ended when people realised that uranium is radioactive, making the pots radioactive too. When nuclear reactors were invented, scientists were able to make nuclear bombs from radioactive plutonium.

There are many radioisotopes which are mainly used in the treatment of hazardous diseases. Cobalt-60 (Co-60) is used in radiotherapy and in food preservation. P-32 and Sr-90 are used to treat skin cancer. Na-24 is used to detect tumors and blood clots and also detect the leakage from underground pipes. Similarly, I-131 is used to investigate the thyroid glands. C-14 is widely used in estimation of the age of ancient artifacts by radiocarbon dating method.

Review Questions

(1) How many electrons, protons and neutrons are present in the following atoms?

$^{16}_{8}O$, $^{39}_{19}K$ and $^{207}_{82}Pb$

(2) Select isotopes and isobars from the following atoms. Give reasons.

$^{39}_{19}K$, $^{238}_{92}U$, $^{27}_{13}Al$, $^{16}_{8}O$, $^{235}_{92}U$, $^{39}_{17}Cl$

Key Terms

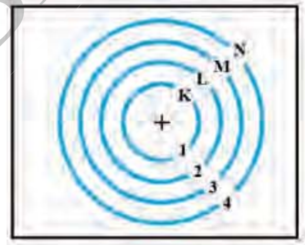
- **Atoms** are the smallest particles into which an element can be divided.
- The number of protons in the nucleus of an atom is known as **atomic number (Z)** of an element. The atomic number never changes.
- The **mass number (A)** of the element is the sum of the number of protons and neutrons or the total number of nucleons in the nucleus of an atom of that element.
- The **nucleon number** of an element is total number of protons and neutrons in the nucleus of its atom.
- Atoms of the same element that have the same number of protons but different number of neutrons are called **isotopes.** They are the atoms of the same element with different masses.
- **Isobars** are the atoms with same mass number but different atomic number.

3.2 ELECTRONIC STRUCTURES (ELECTRON CONFIGURATIONS)

The way in which electrons are arranged around the nucleus of an atom is very important because the electron arrangement determines the chemical properties of the atom. The arrangement of electrons in an atom was suggested in 1913 by Niels Bohr.

(a) Main Shells

Electrons move round the nucleus in definite **shells** or **orbits**. Each shell is numbered 1, 2, 3, 4, and so on, going outwards from the nucleus. They are also known as K shell, L shell, M shell, N shell, and so on, as shown in Figure 3.3. Each shell has different energy levels and can contain a limited number of electrons. In general, the closer the shell is to the nucleus, the lower is its energy. The maximum number of electrons in each shell which could contain can be calculated by the formula $2n^2$, where n is the shell number.



shell number	1	2	3	4
main shell	K	L	M	N
maximum number of electrons	2	8	18	32

Figure 3.3 Diagrammatic Representation of the Nucleus and the Electron Shells in an Atom (not to scale)

(b) Sub-shells

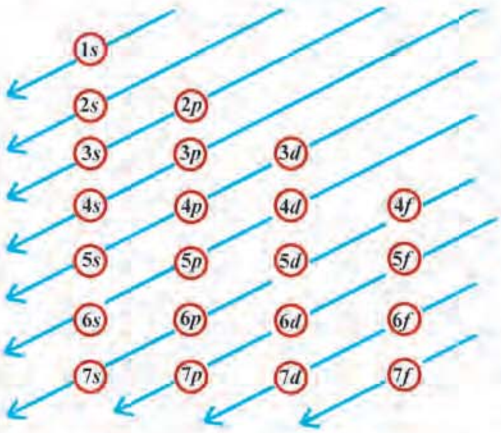


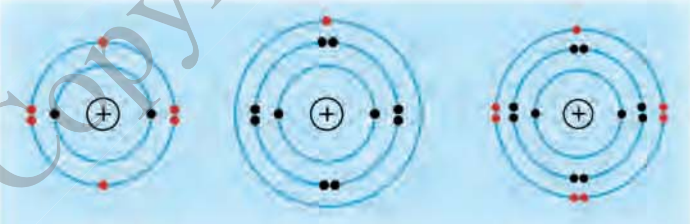
Figure 3.4 Order of Occupancy of Sub-shells

Each main shell is again divided into a number of **sub-shells (orbitals)**, *s*, *p*, *d* and *f*. The letters used for sub-shells notations: *s* stands for sharp; *p* for principal; *d* for diffuse and *f* for fundamental. The maximum number of electrons in *s*, *p*, *d* and *f* sub-shells is 2, 6, 10 and 14, respectively. The shell number 1 (K shell), has only *s* sub-shell, the shell number 2 (L shell) has *s* and *p* sub-shells, the shell number 3 (M shell) *s*, *p* and *d* sub-shells and the shell number 4 (N shell) *s*, *p*, *d* and *f* sub-shells and so on. Therefore, the order of filling the sublevels is given as: $\mathbf{1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6}$**...** .

The order in which the sub-shells fill is shown in the Figure 3.4, which is arranged so that it is easily remembered. One simply lists the sub-shells in order, starting each shell with a new line. The order of filling them is found by crossing them with diagonal arrows.

(c) Arrangement of Electrons

The electronic structures of some elements (arrangement of electrons in an atom of O, Na and Cl) described as integer as well as complete and essential electron configurations are shown in Figure 3.5.



integer form:	O (2.6)	Na (2.8.1)	Cl (2.8.7)
complete electronic structure:	$1s^2\ 2s^2\ 2p^4$	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
essential electronic structure:	$2s^2\ 2p^4$	$3s^1$	$3s^2\ 3p^5$

Figure 3.5 Electronic Structures of Oxygen, Sodium and Chlorine

This is the arrangement of electrons in shells around the nucleus. Each shell can hold a certain maximum number of electrons. The electrons always go into the shell nearest to the nucleus which has lowest energy. Once the shell is filled up, the electrons go into next available shell and so on. Hence, the first shell (K) can hold up two electrons which fill up in *s* sub-shell, subsequent shell (L) up to eight electrons in which two in *s* sub-shell and six in *p* sub-shell and so on.

(d) Valence of the Elements (Combining Capacity)

The shell which is furthest from the nucleus is called the **outer shell**. It is also called the **valence shell**. The electrons in this shell are known as **outer electrons** (or) **valence electrons**. These electrons are involved in chemical reactions.

The number of outermost shell electron ≤ 4, Valence = number of outermost shell electrons

The number of outermost shell electron > 4, Valence = 8 − number of outermost shell electrons

For example,

The complete electronic structure of sodium is $1s^2\,2s^2\,2p^6\,3s^1$.

The essential electronic structure of sodium is $3s^1$.

The outermost shell electron is 1 and valence is 1.

In the case of fluorine, the essential electronic structure is $2s^2\,2p^5$, the outermost shell electron is 7 and its valence is therefore 1, i.e., $8 - 7 = 1$.

Review Question

Use Periodic Table to complete the following table:

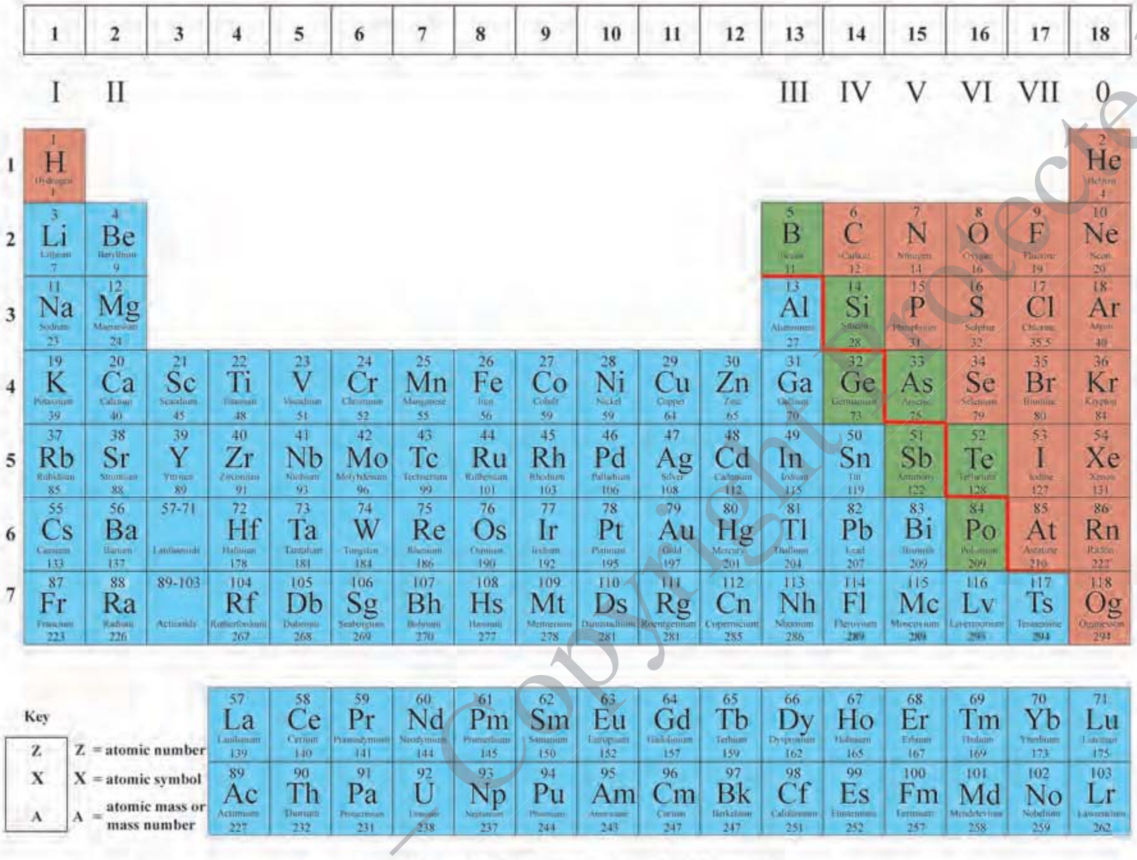
Element	Integer electronic structure	Complete electronic structure	Essential electronic structure	Valence
Li		$1s^2\,2s^1$	$2s^1$	
B		$1s^2\,2s^2\,2p^1$		3
Na	2.8.1		$3s^1$	
Al	2.8.3	$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$		
	2.8.8.1		$4s^1$	
	2.8.7		$3s^2\,3p^5$	

Key Terms

- The distribution of electrons in an atom of an element is called the **electronic structure**.
- The arrangement of all the electrons of an atom of the element in appropriate sub-shells is known as the **complete electronic structure**.
- The representation of the arrangement of valence electrons of an atom of the element in appropriate sub-shells is called the **essential electronic structure**.
- **Valence** is the number of electrons in the outermost shell when the number of electrons in the outermost shell is 4 or less . On the other hand, valence is equal to 8 minus number of electrons in the outermost shell when the number of electrons in the outermost shell is greater than 4.

3.3 THE PERIODIC TABLE

The Periodic Table was devised in 1869 by the Russian chemist Dmitri Mendeleev (1834-1907). His Periodic Table was based on the chemical and physical properties of the 63 elements that had been discovered at that time.



	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	I	II											III	IV	V	VI	VII	0
1	1 H 1																	2 He 4
2	3 Li 7	4 Be 9											5 B 11	6 C 12	7 N 14	8 O 16	9 F 19	10 Ne 20
3	11 Na 23	12 Mg 24											13 Al 27	14 Si 28	15 P 31	16 S 32	17 Cl 35.5	18 Ar 40
4	19 K 39	20 Ca 40	21 Sc 45	22 Ti 48	23 V 51	24 Cr 52	25 Mn 55	26 Fe 56	27 Co 59	28 Ni 59	29 Cu 64	30 Zn 65	31 Ga 70	32 Ge 73	33 As 75	34 Se 79	35 Br 80	36 Kr 84
5	37 Rb 85	38 Sr 88	39 Y 89	40 Zr 91	41 Nb 93	42 Mo 96	43 Tc 99	44 Ru 101	45 Rh 103	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131
6	55 Cs 133	56 Ba 137	57-71 Lanthanides	72 Hf 178	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 Tl 204	82 Pb 207	83 Bi 209	84 Po 209	85 At 210	86 Rn 222
7	87 Fr 223	88 Ra 226	89-103 Actinides	104 Rf 267	105 Db 268	106 Sg 269	107 Bh 270	108 Hs 277	109 Mt 278	110 Ds 281	111 Rg 281	112 Cn 285	113 Nh 286	114 Fl 289	115 Mc 289	116 Lv 293	117 Ts 294	118 Og 294

57 La 139	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm 145	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 162	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
89 Ac 227	90 Th 232	91 Pa 231	92 U 238	93 Np 237	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 262

Key
Z = atomic number
X = atomic symbol
A = atomic mass or mass number

Figure 3.6 The Periodic Table

In the modern Periodic Table the 118 known elements are arranged in order of increasing atomic number (Figure 3.6). This table helps chemists to understand the elements better and to predict properties of elements.

There are 18 vertical columns and 7 horizontal rows in the Periodic Table. The vertical columns are called **groups**. The horizontal rows are called **periods**. Those elements with similar chemical properties are found in the same columns or groups.

Group I elements (except hydrogen) are called the **alkali metals**.

Group II elements are called the **alkaline earth metals**.

Group VII elements are called the **halogens**.

Group 0 elements are known as the **noble gases** or **inert gases**.

Groups I and II consist of *s*-block elements. Groups III, IV, V, VI,VII and group 0 consist of ***p*-block elements**. **Transition elements** are ***d*-block elements**. **Inner transition elements** (lanthanides and actinides series), also known as **rare earth elements** are the ***f*-block elements**.

Chemistry in Society

- Lithium is used to make batteries that power electronic devices like digital cameras.
- Sodium vapour lamps give off yellow-orange light and are often used in street lamps.
- The halogens have many varied uses: fluoride in toothpaste to help reduce dental decay; chlorine in household bleach to kill bacteria, to whiten clothes, and to clean swimming pool; bromine as a fire retardant; and iodine in photographic reproduction.
- The metalloid silicon is used to make silicon chips.
- Helium is used for filling weather or advertisement balloons and airships.
- Neon is used in making light and advertising signs.
- Argon is used to fill tungsten bulbs to provide an inert (unreactive) atmosphere that prevents oxidation of the filament.

Review Question

You are given the following elements: A to G represent unknown elements.

$_{10}A$, $_{17}B$, $_{3}C$, $_{9}D$, $_{11}E$, $_{18}F$, $_{19}G$.

(a) Write down the electronic structures of these elements.
(b) Which elements are alkali metals?
(c) Which elements are noble gases?
(d) Which elements are halogens?
(e) Which elements are *s*-block elements?

Key Terms

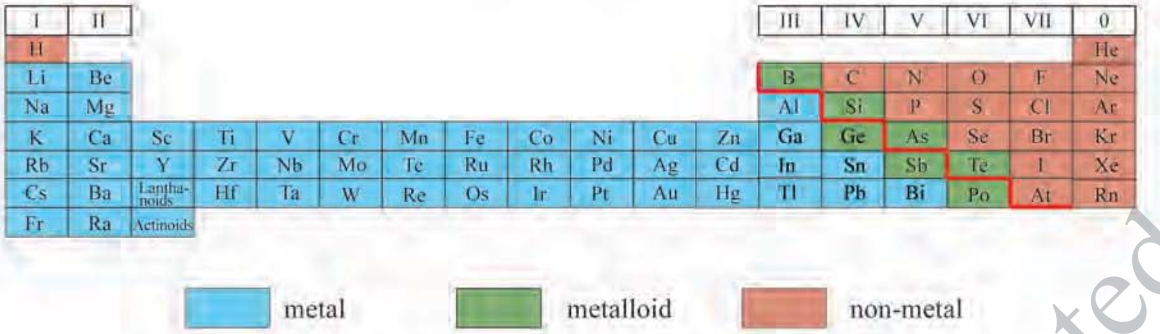
- The **Periodic Table** is a list of chemical elements arranged in order of atomic number in rows, so that elements with similar electronic structures (and hence, similar chemical properties) appear in vertical columns. There are 18 vertical columns and 7 horizontal rows in the Periodic Table.
- The elements are classified as alkali metals (group I), alkaline earth metals (group II), halogens (group VII) and noble gases or inert gases (group 0).

3.4 PERIODIC PROPERTIES

The Periodic Table can be used to predict the properties of elements. Different elements have different periodic properties, such as metallic and non-metallic character, electronegativity, size, ionisation energy and electron affinity of the elements.

(a) Metallic and Non-metallic Character

As we go from left to right across the Periodic Table, the elements change in properties from metals to non-metals. The Periodic Table can be divided into two as shown by the dark line that starts beneath boron (Figure 3.7).



I	II											III	IV	V	VI	VII	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	Lantha-noids	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Actinoids															

🟦 metal 🟩 metalloid 🟥 non-metal

Figure 3.7 Metals, Non-metals and Metalloids

All elements to the left of the red line are metals.

The elements to the right of the red line are non-metals.

For along each side of the red line are elements that have the properties of both metals and non-metals. These elements except Al are called **metalloids**. Hydrogen is also a non-metal.

(b) Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electron; the higher the electronegativity, the greater is an atom's attraction for electrons. In the Periodic Table, in general, electronegativity of elements increases from left to right across a period, and the bottom of a group to the top (although this does not apply to the transition metals).

Atoms with high electronegativity tend to form negative ions (e.g., fluorine, oxygen). Atoms with low electronegativity, which is highly **electropositive elements** (e.g., caesium, potassium) tend to form positive ions. In general, metals are the electropositive elements and non-metals are the **electronegative elements**. The most electronegative element is fluorine, and the least electronegative (most electropositive) element is caesium.

The noble gases or inert gases are neither electropositive nor electronegative. It is because they have very stable electronic structures and have little tendency to gain or lose electrons.

(c) Sizes

Atomic sizes (Atomic radii)

Atomic size is generally described by the radius of an atom. Atomic radii decrease from left to right across a period. This is because ongoing from left to right across a period, the nuclear charge (atomic number) increases while the added electrons enter the outermost shell. The increased nuclear charge attracts the electrons in the outermost shell closer to the nucleus. Hence, the shell contracts resulting in smaller atoms.

The atomic radii increase from top to bottom down a group in the Periodic Table. This is because as the number of electrons increases, these additional electrons are in the larger electron shells make farther and farther from the nucleus. Hence, the shell expands resulting in larger atoms.

Ionic sizes

When one or more electrons are removed from a metal atom, a positive ion (a cation) is formed. Both a positive ion and the parent neutral atom have same nuclear charges. There is a lesser number of electrons in the positive ion. Hence, the repulsion between electrons is reduced in the positive ion. Thus, a positive ion is always smaller than its parent atoms.

When one or more electrons are added to a neutral atom, a negative ion (an anion) is formed. Both a negative ion and the parent neutral atom have same nuclear charges. There are a greater number of electrons in the negative ion. Hence, the repulsion between electrons is increased in the negative ion. Thus, a negative ion is always larger than its parent atom.

(d) Ionisation Energy

The amount of energy required to remove an electron from a gaseous atom to form a gaseous ion is called the **ionisation energy.** These electrons are held strongly within the atom by the attraction of the nucleus. Thus, ionisation process can be expressed in an equation.

$$\text{element(g)} \quad + \quad \text{ionisation energy} \quad \longrightarrow \quad \text{cation(g)} \quad + \quad \text{electron}$$

$$\text{M(g)} \quad + \quad \text{energy} \quad \longrightarrow \quad \text{M}^{+}\text{(g)} \quad + \quad \bar{e}$$

Ionisation energies measure how tightly electrons are bound to atoms. Low ionisation energies indicate ease of removal of electrons. As ionisation energy (I) increases, atoms are harder to ionise. Successive ionisation for electrons are represented by I_i (i = 1, 2, 3,….).

$$I_1 < I_2 < I_3 < \ldots$$

This is because, the nuclear charge increases across a period and the electrons are more strongly held by the force of attraction between the nucleus and the electrons.

In general, ionisation energies increase from left to right across a period. This is because, the nuclear charge increases across a period and the electrons are more strongly held by the force of attraction between the nucleus and the electrons. Therefore, more energy is required to remove an electron from the element.

In general, ionisation energies decrease down a group. This is because the atomic size increases and the outermost electron is farther from the nucleus making it easier to remove it. Therefore, less energy is required to remove an electron from the element.

The noble gases have the highest ionisation energies. This is because the noble gases are known to have the closed electronic structures (the octet) which resist the removal of electrons.

(e) Electron Affinity

The **electron affinity** of an element is the energy released when an electron is added to a gaseous atom to form a gaseous ion.

$$\text{element(g)} \quad + \quad \text{electron} \quad \longrightarrow \quad \text{anion(g)} \quad + \quad \text{energy}$$

$$\text{X(g)} \quad + \quad \bar{e} \quad \longrightarrow \quad \text{X}^{-}\text{(g)} \quad + \quad \text{energy}$$

The electron affinities generally increase from left to right across a period.

The electron affinity decreases on moving down a group. This is because the size of the atom increases and the electron being added goes to higher shells.

Review Questions

(1) From the following groups, select the one which has the largest radius. Give reasons.
(a) Fe^{2+}, Fe^{3+} (b) Cl, Cl^- (c) Li, Na, K (d) C, N, O

(2) Explain which of the following elements has the highest ionisation energy:
(a) 2.8.1 (b) 2.8.5 (c) 2.8.8

(3) Arrange the following elements in order of their increasing electronegativity:
oxygen, carbon, fluorine, nitrogen.

Key Terms

- **Metalloids** are the elements that have the properties of both metals and non-metals.
- Metals are the **electropositive elements**. They tend to lose electrons and form positive ions.
- Non-metals are the **electronegative elements**. They tend to gain electrons and form negative ions.
- The amount of energy required to remove an electron from a gaseous atom to form a gaseous ion is called the **ionisation energy**.
- The **electron affinity** of an element is the energy released when an electron is added to a gaseous atom to form a gaseous ion.

3.5 BONDS BETWEEN ATOMS

Since atoms exist by themselves, they are usually joined together in groups by chemical bonds. There are two main ways of forming chemical bonds between atoms: ionic bonding and covalent bonding. There is another important type of chemical bonding, called metallic bonding. This is only found in metals.

(a) Formation of Bonds

The electronic structures of noble gases are very stable and unreactive, so they do not need to lose or gain any extra electrons to fill up their outermost shells. They do not usually react with other elements to form compounds.

Atoms of most other elements are reactive. They combine with other atom to form molecules or compounds. In forming a chemical bond, atoms gain, lose or share electrons in such a way to attain the stable electronic structures of the noble gases, i.e., to have eight electrons in the outermost shell, which is known as **octet** electron configuration **(Octet Rule)**. Exception is that helium has only two outer electrons. Helium has a **duplet** electron configuration.

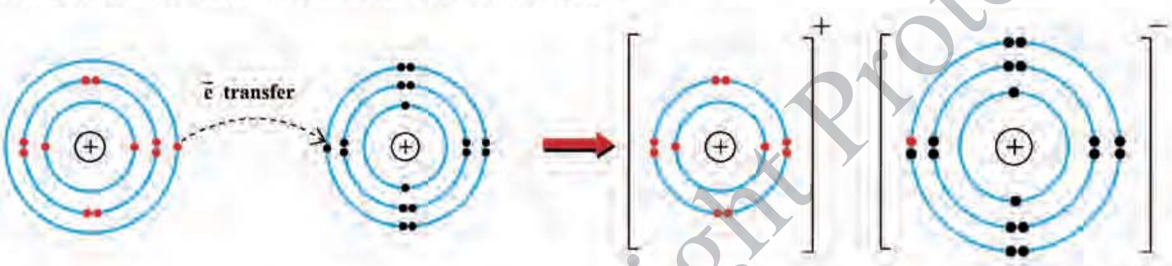
(b) Ions Formation

An ion is formed when an atom loses or gains electrons, so that it has a charge on it. Metals tend to lose electrons to form positively charged ions (cations). Non-metals tend to gain electrons to form negatively charged ions (anions).

(c) Ionic Bond

An ionic bond is formed when there is complete transfer of an electron or electrons from one atom to another resulting in the formation of cations and anions. These oppositely charged ions are held together by a force of electrostatic attraction known as **ionic bond**. This kind of bond occurs mainly between a metal and a non-metal. Compounds that contain ionic bonds are called ionic compounds.

For example, sodium reacts with chlorine to form sodium chloride. In this reaction, the sodium atom loses an electron to become a sodium ion, Na^+. The electron is taken by a chlorine atom to become a chloride ion, Cl^- (Figure 3.8). There is a transfer of an electron from sodium atom to the chlorine atom. In this way, sodium ion and chloride ion achieve the electron configuration of the stable noble gas.



sodium atom (2.8.1) chlorine atom (2.8.7) sodium ion (2.8) chloride ion (2.8.8)

Figure 3.8 Formation of an Ionic Bond in Sodium Chloride

(d) Covalent Bond

A **covalent bond** is formed by the sharing of electron between two atoms by weak intermolecular forces of attraction. Covalent bonds are formed when non-metal reacts with one another.

The bond can be formed between atoms of the same element (e.g., H_2, O_2 and N_2 molecules) or between atoms of the different elements (e.g., CO_2, H_2O, NH_3 and CH_4 molecules).

When a pair of electrons is shared, a single covalent bond is formed. When two pairs of electrons are shared, a double covalent bond is formed.

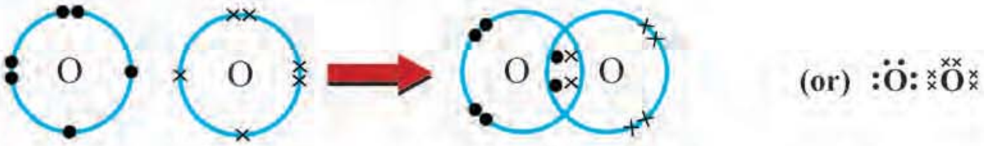
Dot-cross diagrammatic representation is used to explain more explicitly for the bond formation between atoms.

(i) Covalent bond in same elements

H_2 molecule contains a single covalent bond.

two hydrogen atoms one hydrogen molecule

O_2 molecule contains a double covalent bond.



two oxygen atoms one oxygen molecule

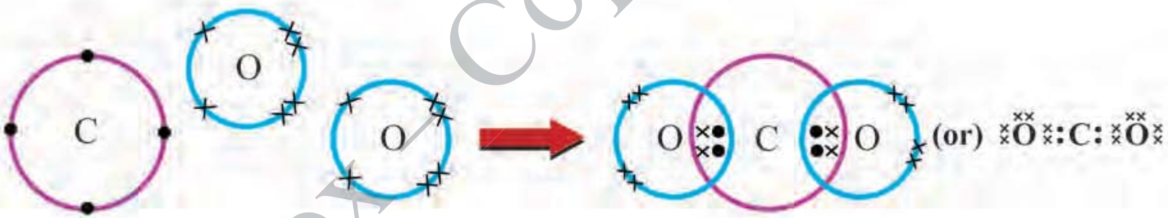
N_2 molecule contains a triple covalent bond.

two nitrogen atoms one nitrogen molecule

(ii) Covalent bond in different elements

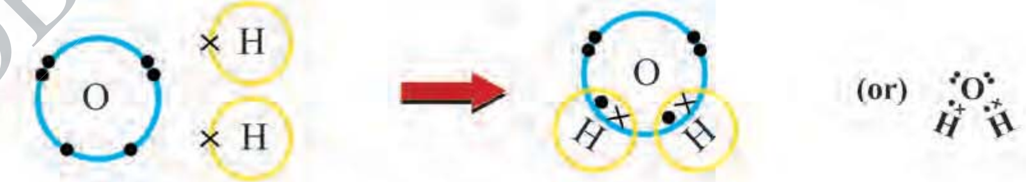
When atoms of different elements are joined together by covalent bonding, a covalent compound or molecular compound is formed.

CO_2 molecule



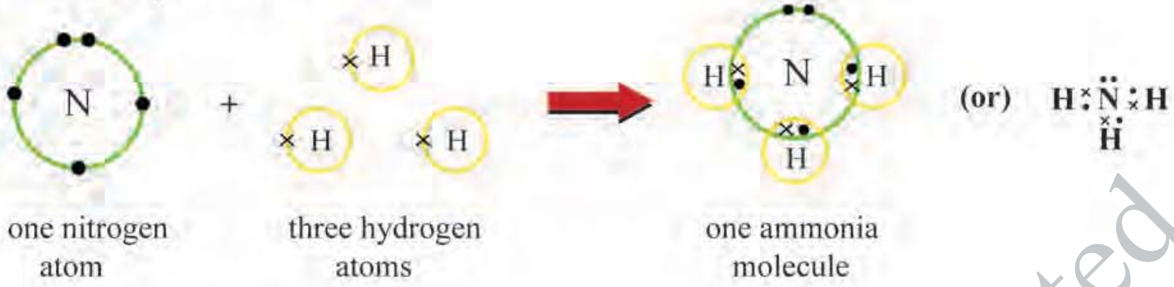
one carbon atom two oxygen atoms one carbon dioxide molecule

H_2O molecule

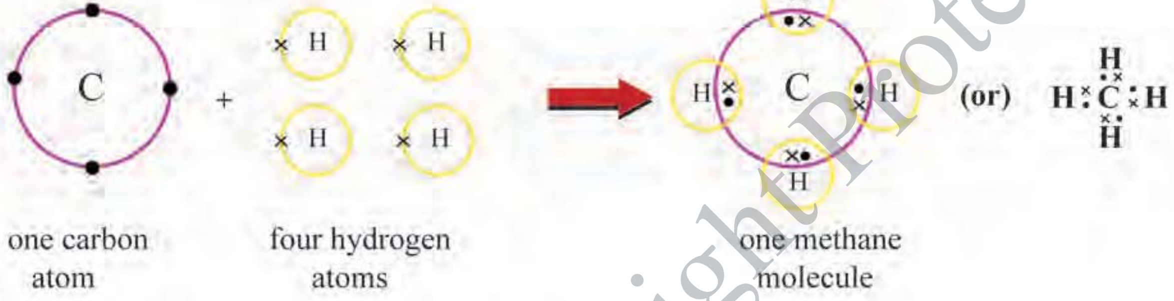


one oxygen atom two hydrogen atoms one water molecule

NH_3 molecule



CH_4 molecule



(e) The Physical Properties of Ionic and Covalent Compounds

The physical properties of compounds depend on the type of bonding in the compounds. The physical properties of ionic and covalent compounds are shown in Table 3.2.

Table 3.2 The Physical Properties of Ionic and Covalent Compounds

No.	Ionic Compounds	Covalent Compounds
1	Ionic compounds do not contain molecules. They consist of aggregates of oppositely charged ions.	Covalent compounds consist of molecules.
2	Ionic compounds are solids and do not vaporise easily.	Simple covalent compounds are gases or volatile liquids (e.g., ammonia, carbon dioxide, ethanol).
3	They conduct electricity when molten or in aqueous solution.	Most of simple covalent compounds do not conduct electricity.
4	Most ionic compounds have high melting points and high boiling points.	Simple covalent compounds have low melting points and low boiling points.
5	Most ionic compounds are soluble in water but not usually soluble in organic solvents such as toluene, ether, benzene, etc.	Simple covalent compounds are usually insoluble in water and soluble in covalent organic solvents, such as toluene, ether, benzene, etc.

Chemistry in Daily Life

Ionic compounds are extremely common in daily life, but not before humans were able to discover, extract and use these compounds. Table salt (NaCl, sodium chloride); baking soda ($NaHCO_3$, sodium hydrogen carbonate); milk of magnesia ($Mg(OH)_2$, magnesium hydroxide) (to treat indigestion); limestone, chalk, marble ($CaCO_3$, calcium carbonate); rust (Fe_2O_3, iron(III) oxide); are examples of ionic compounds. Methane (CH_4, main ingredient in natural gas), hydrochloric acid (HCl), water (H_2O) and ammonia (NH_3) are some examples of covalent compounds.


table salt


baking soda


milk of magneisa

Review Questions

(1) Lithium has the electronic structure 2.1. Fluorine has the electronic structure 2.7. Lithium and fluorine react together to form an ionic lithium fluoride. Draw the arrangement of electrons in fluorine and lithium. Explain how ionic bond is formed in lithium fluoride.

(2) Carbon has the electronic structure 2.4. Chlorine has the electronic structure 2.8.7. Draw the structure of the compound formed between carbon and chlorine.

Key Terms

- A charged particle is an **ion**.
- **Cation** is a positively charged ion.
- **Anion** is a negatively charged ion.
- In forming a chemical bond, atoms gain, lose or share electrons in such a way to attain the stable electronic structures of the noble gases, i.e., to have eight electrons in the outermost shell. This is known as the **octet rule**.
- An **ionic bond** is formed by the complete transfer of an electron or electrons from one atom to another resulting in the formation of cations and anions. These oppositely charged ions are held together by a force of electrostatic attraction known as ionic bond.
- A **covalent bond** is formed by sharing of electrons between two atoms by weak intermolecular force of attraction.

EXERCISES

1. Identify whether each of the following statement is TRUE or FALSE. Give your reasons for considering a statement which is FALSE.
   (a) Atoms of the same element have the same number of neutrons.
   (b) Atoms of different elements can have the same number of nucleon.
   (c) The *s*-subshell contains 2 electrons.
   (d) The 4*s*-sublevel comes before the 3*d* sublevel.
   (e) The fundamental particle not present in a hydrogen atom is the proton.
   (f) The mass of an electron is almost equal to the mass of a proton.

2. Match each of the items given in List A with the appropriate correct item shown in List B.

List A	**List B**
(a) proton	(i) increases down the group
(b) alkali metals	(ii) covalent bond
(c) sharing electrons	(iii) $2 \times n^2$ (n = shell number)
(d) number of electrons in the main nth shell	(iv) lowest electron affinity
(e) atomic size	(v) in the nucleus

3. Fill in the blanks with the correct word(s), phrase, term etc., as necessary.
   (a) Atoms of the same element that have same atomic number but different atomic masses are _____.
   (b) Isobars are the elements with same number of _______.
   (c) From top to bottom in a given group, the atomic number increases and the size of the atom _____.
   (d) Second ionisation energy is _____ than first ionisation energy.
   (e) The number of electrons in outermost shell of $_{11}Na$ is _____ .
   (f) The mass of an atom is mainly concentrated in _____.

4. Select the correct word(s), notation(s), term(s) etc. given in the brackets.
   (a) The atom without neutron(s) is [hydrogen; helium; caesium; lithium].
   (b) In Periodic Table, valence electrons are indicated by [group number; shell number; period number; atomic mass].
   (c) The element at the group VII and the period 3 in the Periodic Table is [X (2.7); Y (2.8.3); Z (2.8.7); Q (2.8.8.2)].
   (d) If number of protons and electrons are 8 respectively, valence electrons are [2; 4; 6; 8] in number.
   (e) [Cl^-; Cl^+; Cl; Cl^{2-}] is stabilised by electron octet.

5. (a) How many electrons, neutrons and protons are there in the following elements? Write down their complete electronic structures.
   $^{12}_{6}C$, $^{56}_{26}Fe$, $^{35}_{17}Cl$, $^{40}_{20}Ca$, $^{48}_{22}Ti$, $^{55}_{25}Mn$, $^{45}_{21}Sc$

   (b) Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A).
   (i) Z = 17, A = 35 (ii) Z = 12, A = 24 (iii) Z = 4, A = 9 (iv) Z=19, A=39

6. (a) Draw diagrammatic representation of the following atoms by using shell or energy level:
   (i) $^{9}_{4}\text{Be}$ (ii) $^{19}_{9}\text{F}$ (iii) $^{28}_{14}\text{Si}$ (iv) $^{39}_{19}\text{K}$
   (b) For the elements in above question, give the valence electrons and the number of neutrons for each.

7. (a) Rewrite the correct complete electronic structures given below.
   (i) $1s^2\ 2s^2\ 2p^4\ 3s^2$ (ii) $1s^2\ 2s^1\ 2p^6$ (iii) $1s^2\ 2s^2\ 2p^6\ 3p^3\ 3s^2$ (iv) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5\ 4s^2$
   (b) What are the atomic numbers of elements whose outermost electrons are represented by (i) $3s^1$ (ii) $2s^2\ 2p^3$ and (iii) $3s^2\ 3p^5$?

8. Which atoms are indicated by the following configurations?
   (a) [He] $2s^1$ (b) [Ne] $3s^2\ 3p^3$ (c) [Ar] $4s^2$

9. (a) Write the complete and essential electronic structures using noble gases as a core for Li, O, Mg, Al, Cl, Ca.
   (b) Give the group, period and valence of the above elements.
   (c) Which of the above element is in period 3 and group II?

10. An element **X** in period 3 of the Periodic Table has six outer shell electrons.
   (a) In which group of the Periodic Table is **X**?
   (b) What is the name and symbol of **X**?
   (c) How many electron shells are there in an atom of **X**?
   (d) Is the element **X** a metal or non-metal?
   (e) What is the atomic number of the element **X**?

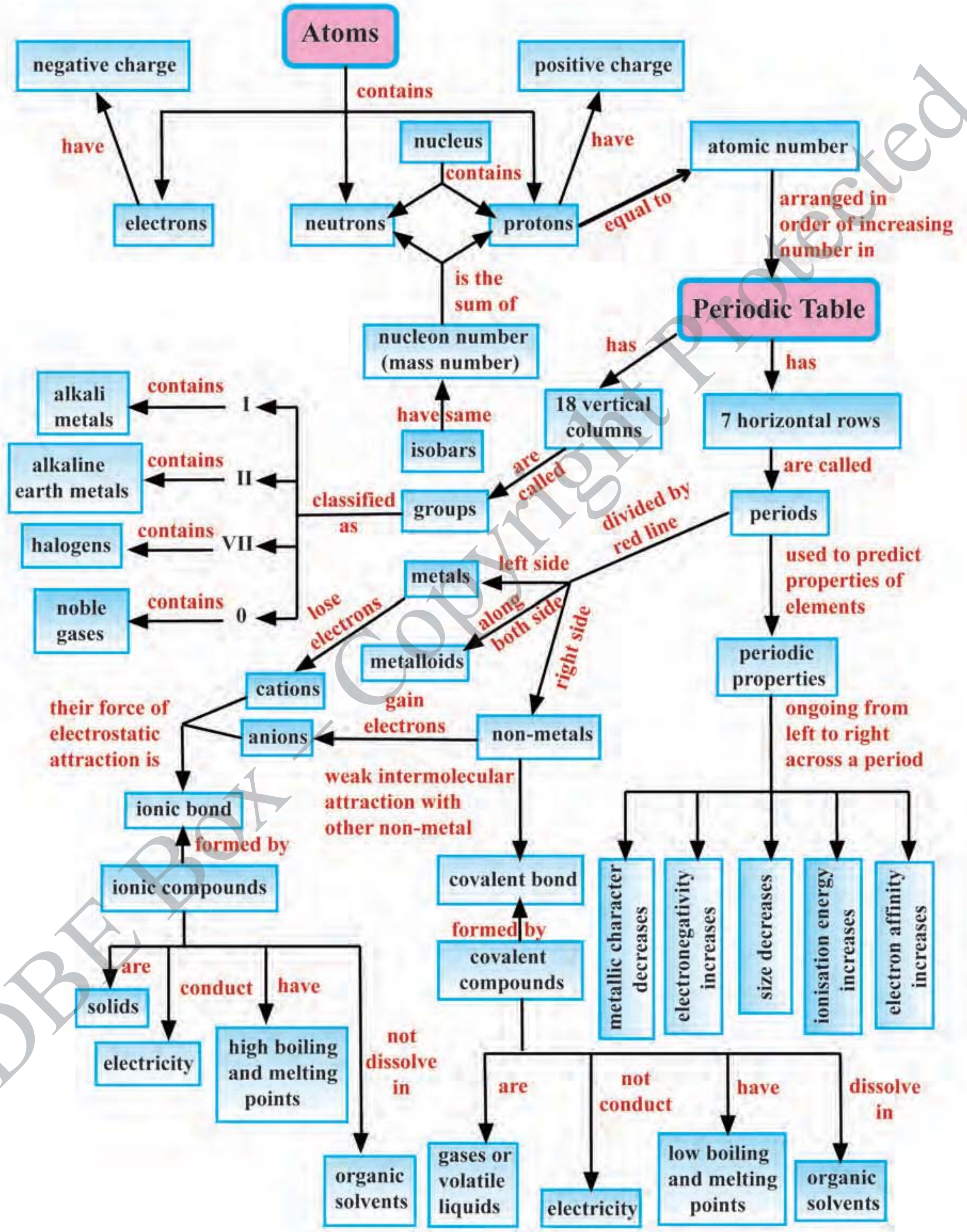
11. Complete the following table and answer the following questions:

	Element **X**	Element **Y**	Element **Z**
atomic number	11	6	-
number of protons	-	-	16
number of neutrons	12	6	16
mass number	-	12	32
electronic structure	2.8.1	-	-
valence	-	4	-
position in Periodic Table	-	-	groupVI, period 3

   (a) Which of the above elements **X**, **Y** and **Z** is a metal?
   (b) What type of bonding exists between **X** and **Z**?
   Write down the most likely formula of this compound using the symbols **X** and **Z**.

CHAPTER REVIEW

(Concept Map)



CHAPTER 4

THE QUANTITIES OF SUBSTANCES: CHEMICAL CALCULATIONS

The state of the art of chemistry lies on the chemical calculations, which are mainly based on the quantities of chemical terms (especially the chemical symbol and the Mole) and other relevant data. In this modern day, chemical ingredients and their percent amounts have to be labelled under rules and regulations. For instance, a farmer must not only know the soil fertiliser symbols, such as N P K, but also the percent amount of soil nutrients. Therefore, chemical calculations have many important uses and applications in chemistry.

Learning Outcomes

After completing this chapter, students will be able to:

- distinguish among the relative atomic mass of sample elements, the relative molecular mass and the relative formula mass of sample substances;
- write chemical symbols, formulae, word equations and chemical equations based on information provided;
- solve calculations based on chemical equations;
- solve the molar volume of a gas based on information provided;
- explain the connection between the mole and Avogadro's constant;
- solve mole calculations and the mole ratio of the reactants and products based on balanced chemical equations.

4.1 RELATIVE MASSES OF ATOMS AND MOLECULES

Atoms and molecules are particles that make up matter. They have mass. For example, the quantity of a cube (1 cm^3) of gold is heavier than a cube (1 cm^3) of aluminium. However, in chemical calculations, chemists use relative masses of atoms and molecules on the basis of a designated standard unit mass, called atomic mass unit (amu).

(a) Relative Atomic Mass

Atoms of different elements have different masses. Therefore, when chemical calculations are performed involving different atoms, we need to know how one atom is heavier or lighter than the other atoms.The mass of a single atom is so small that it is impossible to weigh it directly. To overcome this problem, we then compare this mass of atoms with the mass of the same number of 'standard' atoms. Now, scientists have chosen to use the isotope carbon-12 as the new standard because carbon-12 is the most abundant isotope than carbon-13 and carbon-14.

The mass of an atom is called its **atomic mass**. For the fact that the size of an atom is too small to be weighed practically, **relative atomic mass (A_r)** is therefore used to represent the mass of an atom of an element. Thus, the relative atomic mass of an element is the average mass of one atom of the element compared with one twelfth the mass of one atom of the carbon-12 (^{12}C) isotope whose mass is exactly 12.

$$\text{Relative atomic mass of an element} = \frac{\text{average mass of one atom of the element}}{\frac{1}{12}\text{ of the mass of one atom of carbon-12}}$$

Atomic mass unit (amu)

Atomic mass unit (amu) is precisely one twelfth the mass of one atom of carbon-12. **1 amu equals one twelfth the mass of one atom of C-12 exactly.**

Example 1: A magnesium atom has twice the mass of a ^{12}C atom. What is the relative atomic mass of magnesium? (C = 12)

$$\text{Relative atomic mass of an element} = \frac{\text{average mass of one atom of the element}}{\frac{1}{12}\text{ of the mass of one atom of carbon-12}}$$

$$\text{Relative atomic mass of magnesium} = \frac{2 \times \text{mass of one atom of carbon-12}}{\frac{1}{12} \times \text{mass of one atom of carbon-12}}$$

$$\text{Relative atomic mass of magnesium} = \frac{2 \times 12}{\frac{1}{12} \times 12} = \mathbf{24}$$

Average relative atomic masses

To calculate the relative atomic mass of an element, we must know (a) the isotopic masses of the various isotopes present in the element and (b) the relative abundance of the isotopes in nature.

Example 2: Magnesium consists of three isotopes. One isotope has the relative mass 24 and its relative abundance is 78.6 %. The second isotope has relative mass 25 and its relative abundance is 10.1 %. The third isotope has relative mass 26 and its relative abundance is 11.3 %. Calculate the average relative atomic mass of magnesium.

The average relative atomic mass of magnesium $= (78.6\ \% \times 24) + (10.1\ \% \times 25) + (11.3\ \% \times 26)$

$$= (\frac{78.6}{100} \times 24) + (\frac{10.1}{100} \times 25) + (\frac{11.3}{100} \times 26)$$

$$= 18.86 + 2.53 + 2.94 = \mathbf{24.33}$$

(b) Relative Molecular Mass and Relative Formula Mass

The **relative molecular mass (M_r)** of a molecule is the mass of one molecule of a substance on a scale where the carbon-12 isotopes has a mass of exactly 12 units (12 amu). Thus, the relative molecular mass (M_r) can be found by adding up the relative atomic masses of all the atoms present in the molecule.

Hence, the relative molecular mass (M_r) is defined as the average mass of one molecule of a substance when compared with one twelfth the mass of one atom of carbon-12.

$$\text{Relative molecular mass of a molecule} = \frac{\text{average mass of one molecule of a substance}}{\frac{1}{12}\text{ of the mass of one atom of carbon-12}}$$

Ionic compounds, such as sodium chloride and magnesium oxide consist of ions and not molecules. For ionic compounds, we use the term **relative formula mass** instead of relative molecular mass.

Example 3: What is the relative molecular mass of carbon dioxide, CO_2? The relative atomic masses of carbon and oxygen are 12 and 16, respectively.

Relative molecular mass of carbon dioxide	=	Relative atomic mass of one carbon atom	+	Relative atomic mass of two atoms of oxygen
	=	12	+	2 × 16
	=	12	+	32
	=	**44**		

Example 4: Calculate the formula mass of copper(II) sulphate, $CuSO_4$.
(Given: Cu = 63.5 amu, S = 32 amu, O = 16 amu)

Formula mass of copper(II) sulphate	=	atomic mass of one copper atom	+	atomic mass of one sulphur atom	+	atomic mass of four atoms of oxygen
	=	63.5 amu	+	32 amu	+	4 × 16 amu
	=	63.5 amu	+	32 amu	+	64 amu
	=	**159.5 amu**				

Review Questions

(1) Calculate the relative formula mass of sodium chloride, NaCl. (Na = 23 and Cl = 35.5)

(2) Calculate the relative molecular mass of glucose, $C_6H_{12}O_6$.
(Relative atomic masses: H = 1, C = 12 and O = 16)

Key Terms

- **Relative atomic mass** is the average mass of one atom of that element compared to one twelfth the mass of one atom of carbon-12. **The relative molecular mass** is the sum of the relative atomic masses of all the atoms in the molecule.
- **Relative molecular mass** is the mass of one molecule of a substance compared to one twelfth the mass of one atom of carbon-12.
- **The relative formula mass of an ionic compound** is the sum of the relative atomic masses of all the atoms in the formula.

4.2 CHEMICAL SYMBOLS, FORMULAE, WRITING AND NAMING FORMULAE

Every compound has a formula as well as a name. The formula is made up of the symbols for the elements. Symbols, formulae and equations are bits of chemical shorthand, useful because of a large amount of information which they contain.

(a) Symbols

A symbol is a shorthand notation of an element. The first letter in the name of an element is usually chosen as the symbol of the element. When the names of two or more elements begin with the same letter, the first letter together with another letter in the name of an element is chosen as the symbol of the element. A symbol represents an atom of the element and the mass of an atom.

Symbols of some elements

The symbols of some elements are taken from English names. The symbols of some other elements, which have been known since earlier times, are taken from Latin names (Tables 4.1 and 4.2).

Table 4.1 Symbols of Some Metallic Elements

English name	Latin name	Symbol
Sodium	Natrium	Na
Potassium	Kalium	K
Iron	Ferrum	Fe
Tin	Stannum	Sn
Lead	Plumbum	Pb
Copper	Cuprum	Cu
Mercury	Hydrargyrum	Hg
Silver	Argentum	Ag
Gold	Aurum	Au
Antimony	Stibium	Sb

English name	Symbol
Calcium	Ca
Barium	Ba
Magnesium	Mg
Aluminium	Al
Manganese	Mn
Zinc	Zn
Chromium	Cr
Platinum	Pt

Table 4.2 Symbols of Some Non- Metallic Elements

Name	Symbol
Carbon	C
Sulphur	S
Phosphorus	P
Oxygen	O
Hydrogen	H

Name	Symbol
Nitrogen	N
Chlorine	Cl
Bromine	Br
Iodine	I

(b) Formulae

(i) Molecular formulae of elements

A molecular formula is a shorthand notation representing a molecule or a compound of a substance. It shows the number of atoms contained in a molecule or a compound.

A molecular formula also represents a molecule of the corresponding elements. **For example**, Cl_2 means a molecule of chlorine. $2O_2$ means two molecules of oxygen.

(ii) Molecular formulae of compounds

For those compounds which exist in the form of molecules, a formula represents a molecule as well as the molecular mass of the compound.

A molecule of carbon monoxide contains an atom of carbon and an atom of oxygen. The molecular formula of carbon monoxide is CO. A molecule of ammonia consists of an atom of nitrogen and three atoms of hydrogen. The molecular formula of ammonia is NH_3.

(iii) Formulae for non-molecular compounds

For those compounds which exist in the form of giant structure, a formula represents the simplest unit of the compound.

For example, in sodium chloride the simplest unit consists of an atom of sodium and an atom of chlorine, because sodium and chlorine combine in the atomic ratio 1:1. So, the formula of sodium chloride is NaCl.

Similarly, the formula of magnesium oxide is MgO.

Note: As compounds of giant structure do not have separate molecules, they cannot have molecular mass. So, formula mass is used instead of molecular mass.

(iv) Empirical formula

The first step in identifying a chemical formula is to find out its empirical formula. The empirical formula of a compound is the simplest formula of the compound. It shows

- the types of elements present in the compound;
- the simplest ratio of the different types of the atoms in the compound.

Example 1: When 1.55 g of phosphorus is completely combusted, 3.55 g of an oxide of phosphorus is produced. Deduce the empirical formula of this oxide of phosphorus. (O = 16, P = 31)

	P	O
Step 1 the mass of each element	1.55	3.55 – 1.55 = 2.00
Step 2 divide by relative atomic masses	$\frac{1.55}{31} = 0.05$	$\frac{2.00}{16} = 0.125$
Step 3 divide by the lowest number	$\frac{0.05}{0.05} = 1$	$\frac{0.125}{0.05} = 2.5$
	1 × 2 = 2	2.5 × 2 = 5
Step 4 empirical formula (must be in integers)	P_2O_5	

An empirical formula can also be deduced from data that give the percentage composition by mass of the elements in a compound.

Example 2: A compound of carbon and hydrogen contains 85.7 % of carbon and 14.3 % of hydrogen by mass. Deduce the empirical formula of this hydrocarbon. (H = 1, C = 12)

	C	H
Step 1 the % by mass	85.7	14.3
Step 2 divide by relative atomic masses	$\frac{85.7}{12} = 7.142$	$\frac{14.3}{1} = 14.3$
Step 3 divide by the lowest number	$\frac{7.142}{7.142} = 1$	$\frac{14.3}{7.142} = 2$
Step 4 empirical formula	CH_2	

(v) Calculation involving formulae

Using molecular formula

For some compounds such as magnesium oxide, the empirical formula accurately shows the number of atoms in the compound. However, it is possible that the actual formula differs from the empirical formula. For example, the empirical formula of phosphorus(V) oxide is P_2O_5, while its actual formula is P_4O_{10}.

We call the actual formula of a compound the molecular formula.

When the empirical and molecular formulae of a compound are different, the molecular formula is always a multiple of the empirical formula.

Molecular formula = n × empirical formula, where n is 1, 2, 3, 4….

Example: A compound has the empirical formula CH_2Br. Its relative molecular mass is 188. Deduce the molecular formula of this compound. (Br = 80, C = 12, H = 1)

Step 1 empirical formula mass of $CH_2Br = 12 + (2 \times 1) + 80 = 94$

Step 2 divide the relative molecular mass by the empirical formula mass = $\frac{188}{94} = 2$

Step 3 multiply the number of atoms in the empirical formula by the number in step 2 $= 2 \times CH_2Br$

So molecular formula is $C_2H_4Br_2$.

Empirical and molecular formulae of some compounds are shown in Table 4.3.

Table 4.3 Empirical and Molecular Formulae of Some Compounds

Compound	Empirical formula	Molecular formula
water	H_2O	H_2O
hydrogen peroxide	HO	H_2O_2
sulphur dioxide	SO_2	SO_2
methane	CH_4	CH_4
cyclopropane	CH_2	C_3H_6
butane	C_2H_5	C_4H_{10}

Percentage composition by mass

The formula of a compound and relative atomic masses can be used to calculate the percentage by mass of a particular element in a compound.

$$\text{\% by mass} = \frac{\text{relative atomic mass} \times \text{number of atoms of that element in particular compound}}{\text{relative molecular (formula) mass of a compound}} \times 100$$

Example 1: Calculate the percentage by mass of iron in iron(III) oxide (Fe_2O_3). (Fe = 56, O = 16)

Step 1 Calculate the mass of an element in the compound.

(i) formula of iron(III) oxide = Fe_2O_3

(ii) relative formula mass of Fe_2O_3 = $[(2 \times 56) + (3 \times 16)] = 160$

Step 2 Calculate the percentage by mass of iron in iron(III) oxide.

(iii) $\text{\% mass of iron} = \frac{\text{relative atomic mass of Fe} \times 2}{\text{relative formula mass of } Fe_2O_3} \times 100$

(iv) $= \frac{56 \times 2}{160} \times 100 = 70\ \%$

Example 2: Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3) fertiliser which is used by farmers to increase the yield of crops. (N = 14, H = 1, O = 16)

formula of ammonium nitrate = NH_4NO_3

relative formula mass of NH_4NO_3 = $[14 + (4 \times 1) + 14 + (3 \times 16)] = 80$

$$\text{\% mass of nitrogen} = \frac{\text{relative atomic mass of N} \times 2}{\text{relative formula mass of } NH_4NO_3} \times 100 = \frac{14 \times 2}{80} \times 100 = 35\ \%$$

(c) Writing and Naming Formulae

It seems to be difficult to learn how to write the formulae and the name of a large number of different compounds. But it is not so difficult if we know

(1) the combining capacity of the atoms of different elements and how to use them in formula writing, and

(2) the rules for naming compounds.

Combining capacity or valence

The combining capacity or valence of an element is represented by the number of atoms of hydrogen, chlorine or sodium that combine with one atom of that element. The term 'valence' is also used to express combining capacity.

Different atoms have different combining capacities. Chlorine, for example will combine with sodium, calcium or aluminium to form NaCl, $CaCl_2$ or $AlCl_3$, respectively. So sodium has a combining capacity of 1, calcium a combining capacity of 2 and aluminium a combining capacity of 3 in these compounds.

(i) Fixed combining capacity of certain elements

The element sodium reacts with other elements to form compounds in which it shows electropositive character and a constant combining capacity of 1.

(ii) Variable combining capacities of certain elements

Two different compounds of copper and chlorine are known. They are copper(I) chloride CuCl, and copper(II) chloride, $CuCl_2$. Other examples include $FeCl_2$ and $FeCl_3$; Hg_2O and HgO.

Oxidation number

Oxidation number describes the combining capacity of the element and also indicates the positive and negative nature of its atoms in the compounds. The oxidation number is related to, but not identical with valence or combining capacity.

The use of oxidation numbers

In the compounds formed by the combination of metals with non-metals, the metals always show electropositive character. The combining capacities of the metals in such cases are expressed by using positive oxidation numbers. The combining capacities of the non-metallic elements are expressed by using negative oxidation numbers. For example, in NaCl, the oxidation number of sodium is + 1, and that of chlorine is − 1.

In compounds formed by the combination of one non-metal with another, the more electropositive element is assigned a positive oxidation number and the other is assigned a negative oxidation number. For example, in HCl, hydrogen is assigned an oxidation number of + 1 and chlorine is assigned an oxidation number of − 1.

Combining capacity and common oxidation number of some elements are shown in Table 4.4.

Table 4.4 Combining Capacity and Common Oxidation Number of Some Elements

Name	Combining capacity	Symbol with oxidation number	Name	Combining capacity	Symbol with oxidation number
sodium	1	Na^{1+}	mercury	1, 2	$Hg^{1+, 2+}$
potassium	1	K^{1+}	silver	1	Ag^{1+}
calcium	2	Ca^{2+}	carbon	2, 4	$C^{2+, 4+}$
barium	2	Ba^{2+}	sulphur	2, 4, 6	$S^{2-, 4+, 6+}$
magnesium	2	Mg^{2+}	phosphorus	3,5	$P^{3+, 5+}$
aluminium	3	Al^{3+}	oxygen	2	O^{2-}
manganese	2, 4, 7	$Mn^{2+, 4+, 7+}$	hydrogen	1	H^{1+}
zinc	2	Zn^{2+}	fluorine	1	F^{1-}
iron	2, 3	$Fe^{2+, 3+}$	chlorine	1	Cl^{1-}
tin	2, 4	$Sn^{2+, 4+}$	bromine	1	Br^{1-}
lead	2, 4	$Pb^{2+, 4+}$	iodine	1	I^{1-}
copper	1, 2	$Cu^{1+, 2+}$	nitrogen	1, 2, 3, 4, 5	$N^{3-, 1+, 2+, 3+, 4+, 5+}$

The rules for writing and naming compounds

For convenience we will explain the formula writing and the naming of compounds in the order of : **binary compounds, acids and acid radicals, bases and basic radicals, salts and hydroxides.**

(i) Binary compounds

A binary compound is a compound which contains two elements only. Metallic binary compounds contain a metal, and non metallic binary compounds contain non-metallic elements only.

Order of symbols in a formula: The more electropositive element present in the compound with positive oxidation number is written in front of the symbol of the element with negative oxidation number. One exception is NH_3 in which the symbol of nitrogen with negative oxidation number (N) is written first.

positive oxidation number	**negative oxidation number**
first symbol	**second symbol**

Writing formulae of binary compounds: In the formula of a compound, the algebraic sum of the oxidation numbers must be equal to zero.

The sum of positive oxidation numbers + the sum of negative oxidation numbers = 0

By applying this rule we can find out the number of atoms of each element which should be present in a formula.

Example : When H and S combine with each other what will be the possible formula? The oxidation number of H is + 1. The oxidation number of S is − 2.

$$H^{1+} + H^{1+} + S^{2-} = (H_2S)^0$$

$$\therefore \text{The formula} = H_2S$$

Other examples are:

Na and Cl	$Na^{1+} + Cl^{1-}$	$= (NaCl)^0$	$\therefore$ The formula = $NaCl$
Ca and O	$Ca^{2+} + O^{2-}$	$= (CaO)^0$	$\therefore$ The formula = CaO
Mg and Cl	$Mg^{2+} + Cl^{1-} + Cl^{1-}$	$= (MgCl_2)^0$	$\therefore$ The formula = $MgCl_2$
Al and O	$Al^{3+} + Al^{3+} + O^{2-} + O^{2-} + O^{2-}$	$= (Al_2O_3)^0$	$\therefore$ The formula = Al_2O_3

Naming binary compounds

Compounds in which the first element has a fixed oxidation number

The elements such as H^{1+}, K^{1+}, Na^{1+}, Ca^{2+}, Zn^{2+} etc., have fixed oxidation numbers.

Binary compounds in which the more electropositive element with fixed oxidation number is in the first place of the formula are named thus;

Name of the first element	Name of the second element ending in **- ide**

(Usually the change to **- ide** is in the second syllable of the name.)

Example:	H_2S	Hydrogen sulphide	$BaCl_2$	Barium chloride
	CaO	Calcium oxide	AlN	Aluminium nitride

Compounds in which the first element has a variable oxidation number

The elements such as C, S, Fe, Cu have variable oxidation numbers.

When the more electropositive element in the compound has variable oxidation numbers the name should be given thus:

(1) For the **naming of non-metallic binary compounds**, Greek prefixes (e.g., mono-, di-, tri-, etc.) are used to indicate the number of atoms of each element in the compound. The name of the second element is ended with the syllable **-ide**. In those cases where there is only one atom of the first element, the use of the prefix mono is not necessary.

di, tri, tetra, etc.	Name of the first element	mono, di, tri, tetra, penta,etc.	Name of the second element ending in **-ide**

For example,

N_2O	dinitrogen monoxide (dinitrogen oxide)	NO_2	nitrogen dioxide
NO	nitrogen monoxide (nitrogen oxide)	N_2O_5	dinitrogen pentoxide
N_2O_3	dinitrogen trioxide	CCl_4	carbon tetrachloride

(2) For **naming the metallic binary compounds**, the name of the more electropositive metallic element with variable oxidation number is given first, followed by the Roman Numeral in brackets to state its oxidation number in the compound and the name of the second element ending **-ide**, is added.

Name of the first element	Oxidation number of the first element (in Roman Numeral)	Name of the second element ending in **-ide**

For example,

$FeCl_2$	iron(II) chloride	Cu_2O	copper(I) oxide
$FeCl_3$	iron(III) chloride	CuO	copper(II) oxide
PbO	lead(II) oxide	HgO	mercury(II) oxide

(ii) Acids and acid radicals

The formula of hydrochloric acid is HCl. In HCl only one atom of hydrogen, H, is present. When H is removed from the formula, the remaining part is -Cl. This represents the acid radical of hydrochloric acid.

The formula of sulphuric acid is H_2SO_4 in which there are two H atoms. When one H is removed from the formula, the remaining part is $-HSO_4$. This represents one acid radical of sulphuric acid. In $-HSO_4$ one H atom is still present. When this H atom is also removed, $-HSO_4$ becomes $> SO_4$. This represents another acid radical of sulphuric acid.

In **naming the acids which have hydrogen atoms** in their molecules the name of acid radical **-ic** is changed to **-ide**, e.g., the acid radical of hydrochloric acid is chloride.

-ic acid, [hydrochloric acid (HCl)] $\longrightarrow$ **-ide**, [chloride Cl^-]

In **naming the acids which have oxygen atoms** in their molecules the names of the acid radicals are as follows:

-ous acid, [nitrous acid (HNO_2)] $\longrightarrow$ **-ite**, [nitrite (NO_2^-)]

-ic acid, [nitric acid (HNO_3)] $\longrightarrow$ **-ate**, [nitrate (NO_3^-)]

In **naming the acid radicals containing H**, the word hydrogen is placed before the name of the acid radical. For example, HSO_3^- = hydrogen sulphite

From the above principles, you can derive the formula, the name and the oxidation number of an acid radical from the name and the formula of the corresponding acid. The name and the oxidation number of the acid radical may be derived from the name of acid as shown in Table 4.5.

Table 4.5 Acids and Acid Radicals

Name of acid	Formula of acid	Acid radical	Name of acid radical	Oxidation number	Number of acid radical
hydrochloric acid	HCl	Cl^-	chloride	−1	1
hydrobromic acid	HBr	Br^-	bromide	−1	1
hydriodic acid	HI	I^-	iodide	−1	1
nitrous acid	HNO_2	NO_2^-	nitrite	−1	1
nitric acid	HNO_3	NO_3^-	nitrate	−1	1
chloric acid	$HClO_3$	ClO_3^-	chlorate	−1	1
carbonic acid	H_2CO_3	HCO_3^- CO_3^{2-}	hydrogen carbonate carbonate	−1 −2	2
sulphurous acid	H_2SO_3	HSO_3^- SO_3^{2-}	hydrogen sulphite sulphite	−1 −2	2
sulphuric acid	H_2SO_4	HSO_4^- SO_4^{2-}	hydrogen sulphate sulphate	−1 −2	2
phosphoric acid	H_3PO_4	$H_2PO_4^-$ HPO_4^{2-} PO_4^{3-}	dihydrogen phosphate hydrogen phosphate phosphate	−1 −2 −3	3

(iii) Bases and basic radicals

The ion formed after removal of hydroxide ions (OH^- ions) from a base is called basic radical. For example, when NaOH (base) loses OH^- ion, it forms Na^+ ion which is basic radical. The name and the oxidation number of the basic radical may be derived from the name of base as shown in Table 4.6.

Table 4.6 Bases and Basic Radicals

Name of base	Formula of base	Basic radical	Name of basic radical	Oxidation number	Number of basic radical
sodium hydroxide	NaOH	Na^+	sodium	+ 1	1
potassium hydroxide	KOH	K^+	potassium	+ 1	1
magnesium hydroxide	$Mg(OH)_2$	Mg^{2+}	magnesium	+ 2	1
calcium hydroxide	$Ca(OH)_2$	Ca^{2+}	calcium	+ 2	1
ammonium hydroxide	NH_4OH	NH_4^+	ammonium	+ 1	1

(iv) Salts

Writing the formula of a salt

The formula of a salt consists of two parts. The first part is the metal atom or the ammonium radical. The second part is the acid radical.

Metal atom or NH_4^+	**Acid radical**
First part	Second part

When writing the formula of a salt, the algebraic sum of the oxidation numbers must be equal to zero.

For example, to get the formula of sodium sulphate, we must combine the Na^+ with the SO_4^{2-}. In order to make the sum of the total oxidation number equal to zero, we must combine $2Na^+$ with SO_4^{2-}.

Then the sum of the oxidation numbers $= 2(+1) + (-2) = 0$

$\therefore$ The required formula $= 2Na^+ + SO_4^{2-} = Na_2SO_4$

Other examples are:

(i) $K^{+} + NO_3^{-} = KNO_3$

(ii) $Ca^{2+} + SO_4^{2-} = CaSO_4$

(iii) $Ca^{2+} + HCO_3^{-} + HCO_3^{-} = Ca\,(HCO_3)_2$

(iv) $NH_4^{+} + NH_4^{+} + SO_4^{2-} = (NH_4)_2SO_4$

Naming salts

Naming the salt containing a metal atom with fixed oxidation number or an ammonium radical

The name of the salt begins with the name of the metal or ammonium radical, followed by the name of the acid radical as shown below.

$Ca\,(NO_3)_2$	calcium nitrate
KNO_3	potassium nitrate
$(NH_4)_2SO_4$	ammonium sulphate

Naming the salt in which the metal atom has a variable oxidation number

The name begins with the name of the metal, with Roman Numeral, indicating the oxidation number of the metal atom and followed by the name of the acid radical.

For example:	$FeSO_4$	iron(II) sulphate
	$Fe_2(SO_4)_3$	iron(III) sulphate

Note: The Roman Numeral represents the oxidation number of the metal atom.

(v) Hydroxides

Writing the formula of hydroxide

All hydroxides include one or more - OH radicals of oxidation number − 1 in their formulae. The first part is the metal atom or the ammonium radical. The second part is the - OH radical.

Metal atom or NH_4^{+}	**Hydroxide radical**
First part	Second part

When writing the formula of a salt, the algebraic sum of the oxidation numbers must be equal to zero.

Naming hydroxides

Hydroxides are named in the same way as the naming of salts, but the name of acid radical is replaced by the word hydroxide. For example,

the metal atom with fixed oxidation number	$NaOH$	sodium hydroxide
	$Ca(OH)_2$	calcium hydroxide

the metal atom with variable oxidation number	$Fe(OH)_2$	iron(II) hydroxide
	$Fe(OH)_3$	iron(III) hydroxide
the ammonium radical	NH_4OH	ammonium hydroxide

Chemistry in Society

- An important use of empirical formula calculation is in organic chemistry. Almost every day a new organic compound is either discovered or made in the laboratory. To find the formula of the new substance, a sample is analysed to obtain the mass of percentage composition of each element in the compound. From the data, the empirical formula is then worked out. The relative molecular mass and the molecular formula are also being determined.
- Sodium hydrogen carbonate (baking soda), $NaHCO_3$, is used in the manufacture of some toothpaste and as a raising agent in food production. The purity of this substance can be obtained by measuring how much carbon dioxide is given off.
- Not only chemists need to know about percentage composition, so do farmers. An important element for the growth of plants is nitrogen. If a soil is low in natural nitrogen compounds, plants do not grow as well. Therefore, a farmer can add required amount of artificial nitrogen fertiliser (urea, $CO(NH_2)_2$) to the soil.
- In the food and pharmaceutical industries, it is crucial to know the purity of the products and the formulae. Therefore, they must be labelled under the food and drugs rules and regulations.
- In some commercial products the chemical formulae are used to describe the chemical compounds.

Review Questions

(1) Write the empirical formulae for (a) hydrazine, N_2H_4 (b) octane, C_8H_{18} (c) benzene, C_6H_6 and (d) ammonia, NH_3.

(2) The composition by mass of a hydrocarbon is 10 % hydrogen and 90 % carbon. Deduce the empirical formula of this hydrocarbon. ($C = 12, H = 1$)

(3) Write the formula of each of the following compounds:
magnesium sulphate, potassium carbonate, lead(II) chloride, zinc oxide, ammonium sulphate, aluminium chloride, sulphur trioxide, sodium bromide

Key Terms

- **Empirical formula** shows the simplest whole number ratio of atoms in a compound.
- **Molecular formula** shows the total number of atoms of each element present in one molecule or one formula unit of the compound.
- **Oxidation number** describes the combining capacity of the element and also indicates the positive and negative nature of its atoms in the compounds. The oxidation number is related to, but not identical with valence or combining capacity.

4.3 CHEMICAL EQUATIONS

When carbon is burnt in air, carbon combines with oxygen to form carbon dioxide. This is a chemical change. The process of undergoing a chemical change is called a chemical reaction. A chemical reaction can be represented by a chemical equation. A balanced chemical equation helps us to calculate the right amount of reactant to use in a reaction.

A chemical equation describes:

(1) the reactants, which are the substances that take part in the reaction,
(2) the products, which are the substances that are produced in the reaction,
(3) the rearrangement of the atoms during the reaction, and
(4) the weight relationship of the reactants and the products.

There are two types of chemical equations: **equation in words** and **equation in symbols**.

In an equation in words the names of the reactants are written on the left-hand side of the equation and the names of the products on the right-hand side of the equation.

In the burning of carbon, carbon and oxygen are the reactants and carbon dioxide is the product.

Equation in words: carbon + oxygen ⟶ carbon dioxide

Equation in symbols: $C + O_2 \longrightarrow CO_2$

(i) Steps in writing chemical equations

Let us take the burning of magnesium as an example.

Step 1	Write the word equation for the reaction.	magnesium + oxygen ⟶ magnesium oxide
Step 2	Write in symbols and the formulae of the reactants and the product under the respective names.	$Mg + O_2 \longrightarrow MgO$
Step 3	Balance the equation so that the number of atoms of each element is equal on both sides of the equation.	$2Mg + O_2 \longrightarrow 2MgO$

In Step 3 a complete balanced chemical equation has to abide the **Law of Conservation of Mass**. That is, the total mass of the reactant(s) is equal to the total mass of the product(s). Therefore, the number of the atoms of each element before and after the reaction must also be equal. For this reason it is necessary to write the balanced equation.

(ii) The physical states of the reactants and products

A complete balanced chemical equation must also show the physical states of the reactants and products, whether they are solid, liquid, gaseous, aqueous or in solution. The abbreviations of physical states are written after the corresponding symbols and formulae.

The following is an example of a complete balanced equation including the physical states:

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Note: s = solid; aq = aqueous or in water solution; l = liquid; g = gas

The following abbreviations are usually used in chemical equations:

(1) Δ = heat (2) $\downarrow$ = formation of precipitate

(3) $\uparrow$ = gas evolved (4) $\rightleftharpoons$ = reversible reaction

For example,

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s)\downarrow + NaNO_3(aq) \quad \text{(precipitation reaction)}$$

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s) \quad \text{(combination reaction)}$$

$$S(s) + O_2(g) \xrightarrow{\Delta} SO_2(g) \quad \text{(combustion)}$$

$$NH_4Cl(s) \rightleftharpoons NH_3(g)\uparrow + HCl(g)\uparrow \quad \text{(reversible reaction)}$$

(iii) Writing ionic equations

Most ionic compounds are soluble in water. They exist as ions in aqueous solution. An ionic equation is a simplified chemical equation that shows the reactions involving ions in solution. Let us now see how an ionic equation is written.

When aqueous sodium chloride is added to aqueous silver nitrate, a white precipitate of silver chloride is formed. This reaction can be represented by the following balanced chemical equation:

$$\text{silver nitrate} + \text{sodium chloride} \longrightarrow \text{silver chloride} + \text{sodium nitrate}$$

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s)\downarrow + NaNO_3(aq)$$

Since $AgNO_3$, NaCl and $NaNO_3$ are soluble in water they exist as ions in aqueous solution.

The chemical equation in terms of ions can be written as:

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)\downarrow + Na^+(aq) + NO_3^-(aq)$$

The Na^+ ions and NO_3^- ions are still ions in solution at the end of the reaction. They have not taken part in the chemical reaction. Such ions are called 'spectator ions'.

Since only Ag^+ ions and Cl^- ions have reacted, the equation for the reaction can therefore be simplified as shown below.

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)\downarrow$$

This is the ionic equation for the reaction between aqueous silver nitrate and aqueous sodium chloride.

Review Questions

(1) Write the balanced chemical equations for the following reactions:
   (a) sodium + chlorine $\longrightarrow$ sodium chloride
   (b) sodium + water $\longrightarrow$ sodium hydroxide + hydrogen
   (c) zinc + sulphuric acid $\longrightarrow$ zinc sulphate + hydrogen

(2) When zinc is added to copper(II) sulphate solution, copper and zinc sulphate are formed. Write the balanced ionic equation for this reaction.

Key Terms

- The substances that take part in the reaction are called **reactants**.
- The substances that are produced in the reaction are called **products**.
- **Law of Conservation of Mass** states that the total mass of the reactant(s) is equal to the total mass of the product(s).

4.4 THE MOLE CONCEPT

The relative atomic masses of elements as well as the relative molecular masses of ions or compounds expressed in amu, are not feasible to be used in the weighable scale, especially in grams or other weighable masses. The relative atomic and molecular masses are also composed of microscopic particles and therefore impossible to count individually. The quantitative use of balanced chemical equations so as to calculate the reactants and products need for translating the 'amu' with weighable gram or kilogram, i.e., in the big scale. The gram or kilogram is the measure of countable unit whereas amu is not countable by weighing. For instance, C-12 which has 12 amu becomes 12 g; Mg-24 amu becomes 24 g, and He-4 amu becomes 4 g. It is clear that atomic mass unit (amu) and gram (g) or kilogram (kg) are interrelated. Numerically, the relative atomic and molecular masses are pure numbers.

The chemists' counting unit is named the mole, and it is defined as equal to the number of atoms in exactly 12 g of ^{12}C. This number of particles is called **Avogadro's constant**, which is equal to 6.02×10^{23}.

(a) The Mole and the Avogadro's Constant

We may refer to the mass of a mole of substance as its molar mass (abbreviation M). The unit of molar mass is g mol^{-1}. It is the relative atomic, molecular or formula mass of the substance in grams.

The number of atoms in a mole of atoms is very large: 6.02×10^{23} atoms. This number is called the Avogadro's constant or Avogadro's number (L).

One mole of a substance is the amount of substance that has the same number of particles (atoms, molecules, etc.) as there are atoms in exactly 12 g of ^{12}C.

The Avogadro's constant applies to atoms, molecules, ions and electrons. So in 1 mole of sodium there are 6.02×10^{23} sodium atoms, and in 1 mole of sodium chloride (NaCl) there are 6.02×10^{23} sodium ions and 6.02×10^{23} chloride ions. A mole of chlorine molecules, Cl_2, contains 6.02×10^{23} chlorine molecules but twice as many chlorine atoms, as there are two chlorine atoms in every chlorine molecule.

Amedeo Avogadro (1776-1856) was an Italian scientist who first deduced that equal volumes of gases contain equal number of molecules under the same conditions of temperature and pressure.

(b) Molar Volume of the Gas

One mole of any gas has a volume of 24 dm^3 or 24,000 cm^3 at room temperature and pressure (r.t.p.). This volume is called the **molar volume of a gas**.

Room temperature and pressure are often taken as conditions of 25 °C and 1 atmosphere.

To convert volumes of gases into moles and moles of gases into volumes, the following relationship is used.

$$\text{number of mole of a gas} = \frac{\text{volume of the gas (in dm}^3\text{)}}{\text{molar volume of the gas at r. t. p.}}$$

$$\text{number of mole of a gas} = \frac{\text{volume of the gas (in dm}^3\text{)}}{24\ \text{dm}^3\ \text{mol}^{-1}\ \text{at r. t. p.}}$$

Note: One mole of every gas occupies 22.4 dm^3 at STP (standard temperature, 0 °C or 273 K and standard pressure, 760 mmHg or 1 atmosphere).

Example: Calculate the volume of 0.5 mol of carbon dioxide at room temperature and pressure (r.t.p.).

$$\text{number of moles of } CO_2 = \frac{\text{volume dm}^3 \text{ of } CO_2}{24\ \text{dm}^3\ \text{mol}^{-1}\ \text{at r. t. p.}}$$

$$\begin{aligned}\text{volume of } CO_2 \text{ at r.t.p.} &= \text{number of moles of } CO_2 \times 24\ \text{dm}^3\ \text{mol}^{-1}\\ &= 0.5 \times 24 = 12\ \text{dm}^3 = \mathbf{12{,}000\ cm^3}\end{aligned}$$

(c) Moles and Mass

The Système International (SI) unit for mass is the kilogram. But this is a rather large mass to use for general laboratory work in chemistry. So chemists prefer to use the relative molecular mass or formula mass in grams (1000 g = 1 kg). You can find the number of moles of a substance by using the mass of substance and the relative atomic mass or relative molecular mass or molar mass.

$$\text{number of moles (mol)} = \frac{\text{mass of a substance in gram (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

To find the mass of a substance present in a given number of moles, you need to rearrange the equation:

$$\text{mass of substance (g)} = \text{number of moles (mol)} \times \text{molar mass (g mol}^{-1}\text{)}$$

Example 1: How many moles of sodium chloride are present in 117.0 g of sodium chloride, NaCl? (Na = 23, Cl = 35.5)

$$\text{molar mass of NaCl} = 23.0 + 35.5 = 58.5\ \text{g mol}^{-1}$$

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{117.0\ \text{g}}{58.5\ \text{g mol}^{-1}} = \mathbf{2.0\ mol}$$

Example 2: What mass of sodium hydroxide, NaOH, is present in 0.25 mol of sodium hydroxide? (H = 1, Na = 23, O = 16)

$$\text{molar mass of NaOH} = 23 + 16 + 1 = 40\ \text{g mol}^{-1}$$

$$\begin{aligned}\text{mass} &= \text{number of moles} \times \text{molar mass}\\ &= 0.25\ \text{mol} \times 40\ \text{g mol}^{-1} = \mathbf{10\ g\ NaOH}\end{aligned}$$

Example 3: Calculate the total number of molecules in 7.10 g of chlorine molecule (Cl_2). (Avogadro's constant = 6.02×10^{23} mol⁻¹; Cl = 35.5)

molar mass of $Cl_2 = 2 \times 35.5 = 71$ g mol⁻¹

$$\text{number of moles of chlorine} = \frac{\text{mass of chlorine molecule}}{\text{molar mass of chlorine}} = \frac{7.1 \text{ g}}{71 \text{ g mol}^{-1}} = 0.1 \text{ mol}$$

$$\begin{aligned}\text{number of molecules of } Cl_2 &= \text{number of moles of chlorine} \times 6.02 \times 10^{23} \text{ molecules mol}^{-1} \\ &= 0.1 \text{ mol} \times 6.02 \times 10^{23} \text{ molecules mol}^{-1} \\ &= \mathbf{6.02 \times 10^{22} \text{ molecules}}\end{aligned}$$

(d) Mole Calculations

(i) Calculations from equations (Reacting masses)

Example 1: Calculate the mass of water produced from the complete combustion of 0.25 mol of methane.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Step1 Write the balanced equation.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Step 2 From the equation, find the ratio of the number of moles of H_2O to the number of moles of CH_4.

$$\frac{\text{number of moles of } H_2O \text{ produced}}{\text{number of moles of } CH_4 \text{ reacted}} = \frac{2}{1}$$

Step 3 Use the ratio to find the number of moles of H_2O produced when 0.25 mole of CH_4 is burnt.

$$\begin{aligned}\text{number of moles of } H_2O &= 2 \times \text{number of moles of } CH_4 \\ &= 2 \times 0.25 \text{ mol} = 0.5 \text{ mol}\end{aligned}$$

Step 4 Multiply the number of moles by the molar mass of H_2O to obtain the mass of H_2O in grams.

$$\begin{aligned}\text{molar mass of } H_2O &= (2 \times 1) + 16 = 18 \text{ g mol}^{-1} \\ \text{mass of } H_2O \text{ in grams} &= \text{number of moles} \times \text{molar mass of } H_2O \\ &= 0.5 \text{ mol} \times 18 \text{ g mol}^{-1} = \mathbf{9 \text{ g}}\end{aligned}$$

(ii) Calculations from equations (Reacting masses and volumes)

Example 2: Magnesium reacts with hydrochloric acid according to the equation:

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Calculate the volume of hydrogen gas, measured at room conditions, produced from the reaction of 14.6 g of hydrochloric acid.

Step 1 Change the mass of HCl into moles.

molar mass of HCl = $1 + 35.5 = 36.5\ \text{g mol}^{-1}$

$$\text{number of moles of HCl} = \frac{\text{mass of HCl (g)}}{\text{molar mass of HCl (g mol}^{-1})} = \frac{14.6}{36.5} = 0.4\ \text{mol}$$

Step 2 Write the chemical equation.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Step 3 From the equation, find the ratio of the number of moles of H_2 to the number of moles of HCl.

$$\frac{\text{number of moles of } H_2}{\text{number of moles of HCl}} = \frac{1}{2}$$

Step 4 Use the ratio to find the number of moles of H_2 produced when 0.4 mole of HCl reacts.

$$\text{number of moles of } H_2 = \frac{1}{2} \times \text{number of moles of HCl} = \frac{1}{2} \times 0.4\ \text{mol} = 0.2\ \text{mol}$$

Step 5 Multiply the number of moles of H_2 gas by the molar gas volume. This gives the volume of H_2 gas produced.

volume of H_2 gas = number of moles × molar gas volume = $0.2 \times 24\ \text{dm}^3$ = **$4.8\ \text{dm}^3$**

Chemistry in Society

- **Medicine**: In order to make drug from its ingredients, someone has to figure out how much of each ingredient is needed to react together to make the final drug. That would have involved using the concept of moles.
- **Plastic**: Some plastics are made from other chemicals, someone has to figure out how much of each ingredient is needed to use, and that would have involved moles.
- **Combustion**: You need to use mole in combustion to know how much air is needed, how much exhaust would be produced, as well as how much heat is evolved.
- **Batteries**: Chemicals in batteries react to produce electricity. People have to figure out how much of each type of chemical is needed to put together in a battery to make its function properly. They would also need to know how much the amount of moles of each reactant is needed.

Review Questions

(1) Calculate the amount of moles in 10.7 g of sulphur atoms. (S = 32)

(2) What is the mass of 0.20 mol of carbon dioxide, CO_2? (C = 12, O = 16)

(3) You have a 56 g sample of iron(II) sulphide, FeS.

(a) How many moles of FeS are there in the sample?

(b) How many molecules of FeS are there in the sample?

(Fe = 56, S = 32, Avogadro's constant = $6.02 \times 10^{23}\ \text{mol}^{-1}$).

(4) How many moles are present in the following volumes of gases at r.t.p.?
  (a) 1.2 dm^3 of sulphur dioxide (SO_2)   (b) 0.24 dm^3 of methane (CH_4)
  (c) 120 cm^3 of carbon dioxide (CO_2)

Key Terms

- **One mole** of a substance is the amount of substance that has the same number of particles (atoms, molecules, etc.) as there are atoms in exactly 12 g of ^{12}C.
- The mass of one mole of a substance is called the **molar mass**.
- Equal volumes of gases contain equal number of molecules at the same temperature and pressure.
- The Avogadro's constant (**Avogadro's number** = 6.02×10^{23}) is the number of entities or a stated type of particles (atoms, ions or molecules) in a mole of those substances.
- One mole of any gas has a volume of 24 dm^3 or 24,000 cm^3 at room temperature and pressure (r.t.p.). This volume is called the **molar volume of a gas**.
- One mole of every gas occupies 22.4 dm^3 at **STP** (standard temperature, 0 °C or 273 K and standard pressure, 760 mmHg or 1 atmosphere).

EXERCISES

1. Match each of the items given in List A with the appropriate item given in List B.

List A	List B
(a) Number of acid radicals in H_2SO_4	(i) 24 dm^3 at r.t.p.
(b) The mass of a compound of giant structure	(ii) MgO
(c) The formula of magnesium oxide	(iii) molar mass
(d) Molar volume of gas	(iv) formula mass
(e) The mass of a mole of substance	(v) 2

2. Calculate the empirical formula of each of the following compounds:
   (a) A compound in which 3 g of carbon combines with 4 g of oxygen.
   (b) Iron oxide in which the weight of iron is 77.7 % and that of oxygen is 22.3 %.
   (c) Water in which hydrogen and oxygen combine in the proportion of 1:8 by weight.
   (Fe = 56, C = 12, O = 16, H = 1)

3. The empirical formulae and relative molecular masses of three compounds, A, B and C are shown as follows. Calculate the molecular formula of each of these compounds. (C = 12, Cl = 35.5, H = 1)

Compound	Empirical formula	Relative molecular mass
A	C_3H_5	82
B	CCl_3	237
C	CH_2	112

4. Hydrocarbons are compounds of carbon and hydrogen only. Hydrocarbon Z is composed of 80 % carbon and 20 % hydrogen.

(a) Calculate the empirical formula of hydrocarbon Z. (C = 12, H = 1)
(b) The relative molecular mass of hydrocarbon Z is 30. Deduce the molecular formula of this hydrocarbon.

5. Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: 39.9 % carbon, 6.7 % hydrogen and 53.4 % oxygen.
   (a) Determine the empirical formula of acetic acid.
   (b) Determine the molecular formula of acetic acid if the molecular mass of acetic acid is 60 amu. (C = 12, H = 1, O = 16)

6. Hydrogen peroxide decomposes according to the equation:

$$2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$$

Calculate the volume of oxygen gas produced at r.t.p. when 1.7 g of H_2O_2 is decomposed. (H = 1, O = 16)

7. Tin(IV) oxide is reduced to tin by heating with carbon. Carbon monoxide is also formed.

$$SnO_2(s) + 2C(s) \xrightarrow{\Delta} Sn(s) + 2CO(g)$$

Calculate the mass of carbon that exactly reacts with 14 g of tin(IV) oxide.
(C = 12, O = 16, Sn = 119)

8. A conical flask contains 68.4 g of octane (C_8H_{18}). How many molecules of octane are there in the flask? (Avogadro's constant = 6.02×10^{23} mol^{-1}; C = 12, H = 1).

9. Calculate the number of atoms in 4 g of bromine molecule (Br_2).
(Avogadro's constant = 6.02×10^{23} mol^{-1}; Br = 80).

10. Solid sodium carbonate reacts with aqueous hydrochloric acid to form aqueous sodium chloride, carbon dioxide and water.

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$$

   (a) Rewrite this equation including state symbols.
   (b) Calculate the number of moles of hydrochloric acid required to react exactly with 4.15 g of sodium carbonate. (C = 12, Na = 23, O = 16)

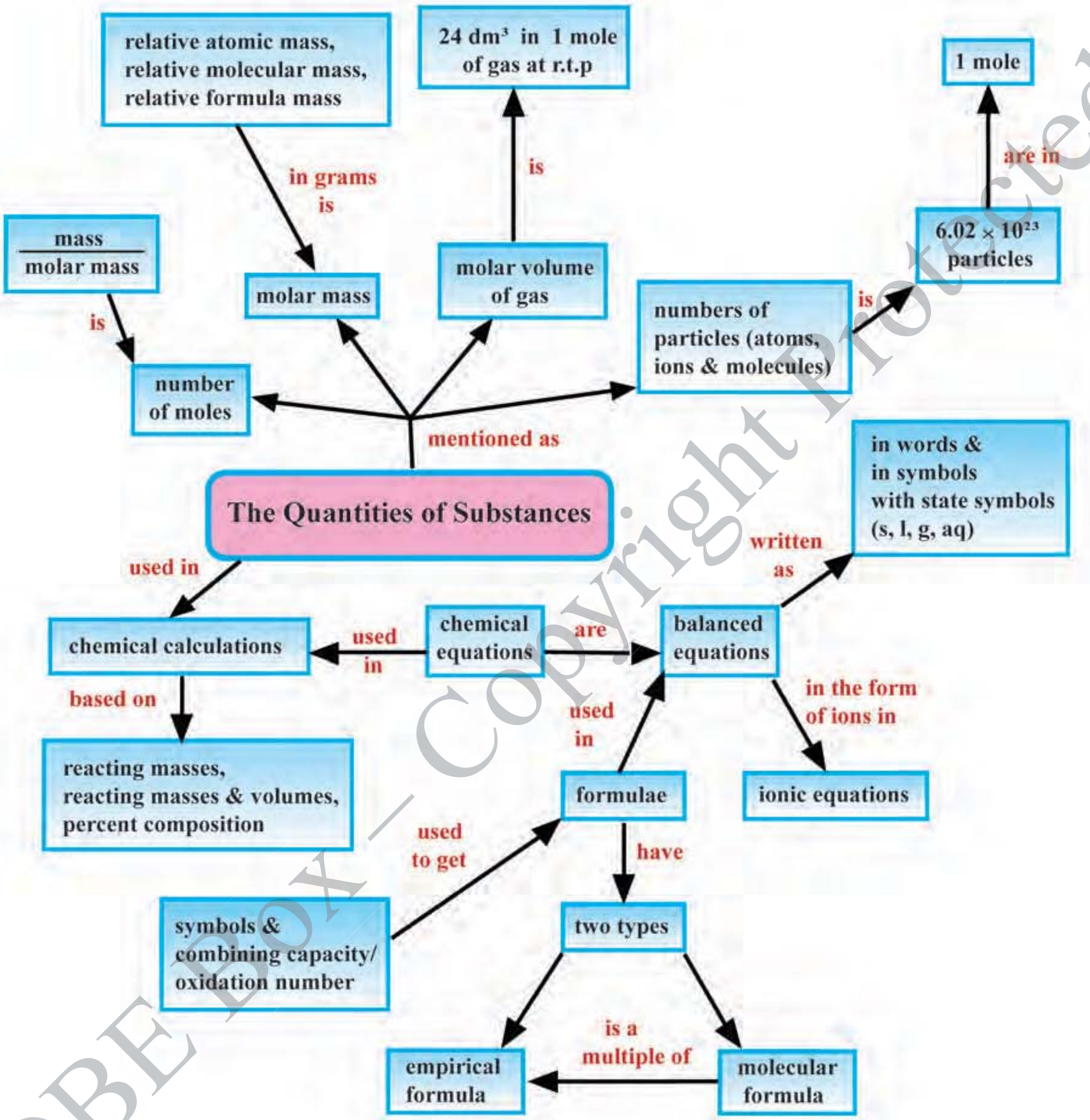
11. Identify the spectator ions and write the ionic equations for the following reactions:
   (a) $Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$
   (b) $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
   (c) $(NH_4)_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l) + 2NH_3(g)$

12. Calculate the amount of substance in moles in each of the following:
   (a) 64.2 g of sulphur molecules (S_8)
   (b) 60.45 g of anhydrous iron(III) nitrate, $Fe(NO_3)_3$. (Fe = 56, N = 14, O = 16, S = 32)

13. Calculate the mass in grams of the following:
   (a) 0.050 moles of sodium carbonate, Na_2CO_3
   (b) 5.00 moles of iron(II) hydroxide, $Fe(OH)_2$
   (C = 12, Fe = 56, H = 1, O = 16, Na = 23)

CHAPTER REVIEW

(Concept Map)



CHAPTER 5

NON-METALS: OXYGEN, CARBON AND HALOGENS

About 50 % of the mass of the Earth's crust consists of oxygen. In combination with carbon, oxygen, hydrogen and nitrogen occur in a large part of plants and animals. Living organisms are mostly made of non-metals. Roughly 96 % of the mass of the human body is made up of just four elements: oxygen, carbon, hydrogen and nitrogen. In this chapter, properties and uses of three non-metals such as oxygen, carbon and halogens (fluorine, chlorine, bromine and iodine) are studied.

You have already learned in Chapter 3 that oxygen with symbol O is the eighth element of the Periodic Table found in the group VI and period 2. The atomic number of oxygen is 8; the mass number is 16. Therefore, the symbol for oxygen atom is $^{16}_{8}O$. Gaseous oxygen molecule is written as O_2.

Carbon is a chemical element with symbol C and atomic number 6, the mass number 12 in group IV and period 2 of the Periodic Table. The symbol is $^{12}_{6}C$. Today, C-12 (exactly) is the standard representative definition of atomic mass unit of all the elements in the Periodic Table.

Halogens are in group VII in the Periodic Table. It consists of five elements: fluorine, chlorine, bromine, iodine and the radioactive element astatine. Halogens are the most reactive non-metals. They react with most metals to form salts. 'Halogen' means **salt-former** in Greek. The molecular formulae are written as F_2, Cl_2, Br_2 and I_2.

Learning Outcomes

After completing this chapter, students will be able to:
- describe the properties and behaviours of oxygen and oxides;
- classify the main types of oxides based on their properties;
- describe the properties and behaviours of carbon;
- explain the allotropy and allotropes of carbon;
- distinguish and compare the properties and behaviours of halogens and halides;
- recognise the role of oxygen, oxides, carbon and halogens in daily life.


Oxygen cylinder


Carbon (Diamond)


Chlorine $Cl_2(g)$


Bromine $Br_2(l)$


Iodine $I_2(s)$

5.1 OXYGEN

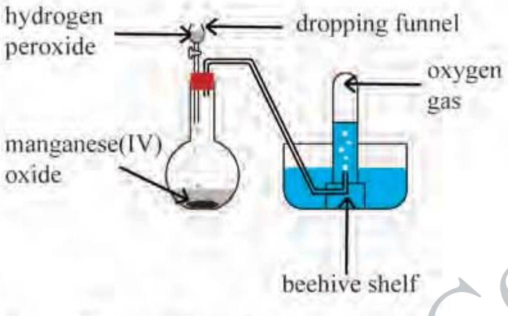
Oxygen is the most abundant element in the Earth's crust. The free element accounts for 21 % of the volume of the atmosphere. Oxygen in the combined state exists in water, sand or silica, silicates and rocks.

(a) Preparation of Oxygen

Oxygen preparation can be demonstrated in a number of ways in the classroom or in the laboratory. Chlorates are principally toxic by injection and inhalation. Therefore, potassium chlorate should not be used for the preparation of oxygen in the laboratory.

Activity (1): Preparation of oxygen from hydrogen peroxide

In the laboratory or in the classroom, oxygen gas can be prepared by using an environmentally friendly liquid hydrogen peroxide of appropriate strength. In this reaction manganese(IV) oxide is used as a catalyst to speed up the chemical reaction. Oxygen gas is collected by the downward displacement of water.



$$\text{hydrogen peroxide} \xrightarrow{MnO_2} \text{water} + \text{oxygen}$$

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

(b) Properties of Oxygen

Physical properties

Oxygen is a colourless gas, without taste or smell. Oxygen is only slightly soluble in water and has about the same relative vapour density as air. Oxygen will not burn, however, it supports combustion. It rekindles any glowing thing (splinter) with a bright flame.

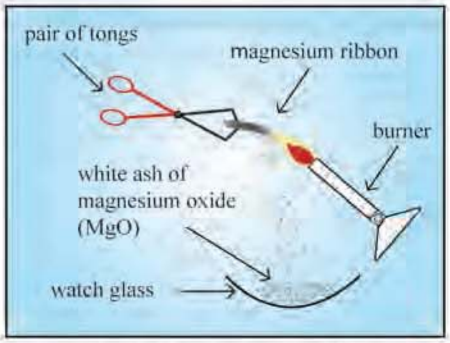
Chemical properties

(i) Action with metals

Metals such as magnesium, iron, copper and zinc react with oxygen forming oxides. A piece of clean magnesium ribbon continues burning in oxygen from air with a dazzling white flame, leaving a white powder as residue. This residue is magnesium oxide.

$$\text{magnesium} + \text{oxygen} \xrightarrow{\Delta} \text{magnesium oxide}$$

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



Burning of magnesium ribbon

Iron slowly becomes oxidised in the presence of air and water, to form hydrated iron(III) oxide, i.e., iron has become rusted.

iron + oxygen + water ⟶ hydrated iron(III) oxide (rust)

$4Fe(s) + 3O_2(g) + 2nH_2O(l) \longrightarrow 2[Fe_2O_3.nH_2O](s)$

Most useful metals or sheets become covered with thin films of metal oxide.

(ii) Action with non-metals

Non-metals such as phosphorus, sulphur and carbon react with oxygen forming oxides. A small piece of phosphorus (only red phosphorus may be used) burns in oxygen giving off white fumes which consist of oxides of phosphorus.

Burning red phosphorus in air

phosphorus + oxygen ⟶ phosphorus(V) oxide

$4P(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$

phosphorus + oxygen ⟶ phosphorus(III) oxide

$4P(s) + 3O_2(g) \longrightarrow P_4O_6(s)$

(c) Oxides

An oxide is a compound containing oxygen and another element. There are six main types of oxides.

(i) acidic oxides (ii) basic oxides (iii) amphoteric oxides
(iv) neutral oxides (v) peroxides (vi) compound oxides

(i) Acidic oxides

An acidic oxide is an oxide of non-metal. It reacts with basic oxides to give salts only.

carbon dioxide + sodium oxide ⟶ sodium carbonate

$CO_2(g) + Na_2O(s) \longrightarrow Na_2CO_3(s)$

Acidic oxides react with alkali solutions to give salts and water.

Test for carbon dioxide

carbon dioxide + calcium hydroxide ⟶ calcium carbonate + water

$CO_2(g) + Ca(OH)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l)$

Some acidic oxides are soluble in water but some are not. CO_2, SO_2, SO_3 and P_4O_{10} are soluble oxides while SiO_2 is an insoluble oxide. Soluble acidic oxides dissolve in water to form acidic solutions. These solutions turn blue litmus red.

sulphur trioxide + water ⟶ sulphuric acid

$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$

(ii) Basic oxides

A basic oxide is an oxide of metal. A basic oxide reacts with an acid to produce a salt and water only. It neutralises acids.

sodium oxide + hydrochloric acid (dil.) → sodium chloride + water

$$Na_2O(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l)$$

Some basic oxides are soluble in water but some are not. Na_2O and K_2O are soluble oxides while MgO, CuO and Ag_2O are not. Some basic oxides react with water forming hydroxide solutions (alkalis). These solutions turn red litmus blue.

sodium oxide + water → sodium hydroxide

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

(iii) Amphoteric oxides

An amphoteric oxide is a metallic oxide which possesses both basic and acidic properties. It reacts with both acids and alkalis to form salt and water. (e.g., ZnO, Al_2O_3, PbO)

zinc oxide + sulphuric acid (dil.) → zinc sulphate + water

$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$$

zinc oxide + sodium hydroxide → sodium zincate + water

$$ZnO(s) + 2NaOH(aq) \longrightarrow Na_2ZnO_2(aq) + H_2O(l)$$

(iv) Neutral oxides

The neutral oxide is an oxide which shows neither basic nor acidic character. (e.g., CO, N_2O)

(v) Peroxides

Those oxides that react with an acid to give salt and hydrogen peroxide are called peroxides. (e.g., BaO_2, Na_2O_2)

barium peroxide + sulphuric acid (dil.) → barium sulphate + hydrogen peroxide

$$BaO_2(s) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + H_2O_2(aq)$$

PbO_2, MnO_2 and NO_2 are not peroxides. They do not give hydrogen peroxides on reaction with acids.

(vi) Compound oxides

A compound oxide is an oxide, formed by the combination of two different oxides of the same element. (e.g. Pb_3O_4, Mn_3O_4, Fe_3O_4)

Pb_3O_4 is a compound oxide consisting of lead(II) and lead(IV) oxides. It can be written as di-lead(II) lead(IV) oxide or red lead oxide, ($2PbO.PbO_2$). When red lead oxide reacts with dilute nitric acid, lead(II) oxide only reacts with the acid to form lead(II) nitrate.

red lead oxide + nitric acid (dil.) ⟶ lead(II) nitrate + water + lead(IV) oxide

$$Pb_3O_4(s) + 4HNO_3(aq) \longrightarrow 2Pb(NO_3)_2(aq) + 2H_2O(l) + PbO_2(s)$$

Chemistry in Society

- Man and animals use oxygen from the air in respiration. Oxygen is used to help patients with breathing difficulties. Mountaineers and under water divers need oxygen cylinders with them. In aquatic habitat, organisms use oxygen dissolved in water.
- In living organisms, the oxygen intake is used for the breakdown of the glucose molecules to produce energy.
- Oxygen is used in steel work, and oxyacetylene gas mixture is used for cutting and welding metals.
- Liquid oxygen is carried on rockets to support the fuel burns. Oxygen is essential in combustion processes such as the burning of fuels.

Usefulness of some oxides in society is described in the following Table:

Oxides	Uses	Oxides	Uses
CO_2	fire extinguishers	SiO_2	manufacture of glass
SO_2	bleaching agent	Pb_3O_4	pigment used in paints
MgO	laxative	ZnO	skin conditioners, cosmetics
CaO	cement production	H_2O_2	antiseptic, hair dyes and toothpaste

Review Questions

(1) The mountaineers and under water divers need to carry oxygen cylinders. Why?

(2) Why manganese(IV) oxide is used as a catalyst for the preparation of oxygen in the laboratory?

(3) Identify the class of oxides to which each of the following belongs:

(a) carbon monoxide (b) red lead oxide (c) sulphur dioxide
(d) sodium peroxide (e) copper(II) oxide (f) lead(II) oxide

Key Terms

- An **acidic oxide** is a non-metallic oxide which reacts with basic oxide to produce salt.
- A **basic oxide** is a metallic oxide which reacts with acid to form salt and water.
- An **amphoteric oxide** is a metallic oxide which possesses both basic and acidic properties.
- A **neutral oxide** does not react with either acids or bases.
- A **peroxide** reacts with an acid to produce salt and hydrogen peroxide.
- A **compound oxide** is the combination of two different oxides of the same element.

5.2 CARBON

Carbon is found in nature as diamond and graphite. Fullerene and graphene are synthetic carbon. Carbon can also be found as compounds combining with other elements in petroleum, coal, natural gas, limestone, carbon dioxide and sugar ($C_{12}H_{22}O_{11}$), etc. In addition, all living things have carbon containing compounds such as carbohydrates, fats, proteins and nucleic acids, etc.

(a) Allotropy and Allotropes of Carbon

If an element, can exist more than one form, in the same physical state, it is said to exhibit allotropy or polymorphism. The different forms of an element in the same physical state that possess different physical properties are known as allotropes of that element. They may have different chemical properties. For example,

Diamond, graphite, fullerene and graphene are allotropes of carbon.
Oxygen and ozone are allotropes of oxygen.
Rhombic sulphur and monoclinic sulphur are allotropes of sulphur.

Diamond

In diamond, each carbon atom is surrounded by four other carbon atoms (Figure 5.1). It has a giant structure. It contains millions of carbon atoms in a three dimensional network of strong carbon-carbon covalent bonds. Therefore, it is very hard and has a very high melting point (3550 °C). Diamond is the hardest among all naturally occurring substances. It is transparent and shines in presence of light.

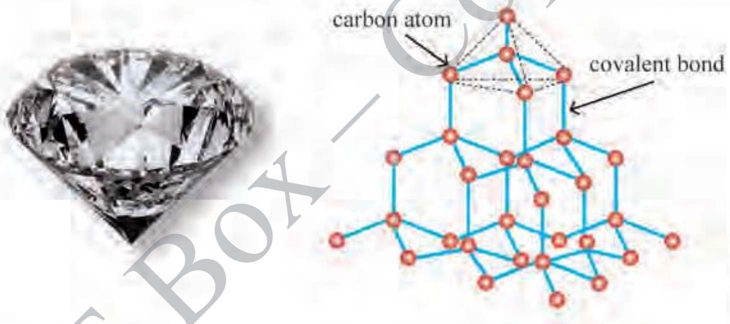


Figure 5.1 Diamond and Its Structure

Graphite

In graphite, each carbon atom is surrounded by three other carbon atoms in the same plane, and therefore layers of hexagons are obtained (Figure 5.2). The distance between the layers is more than the distance between adjacent carbon atoms and so the layers are weakly bonded to each other. Therefore, graphite is soft.

Due to its layered structure, graphite is soft and has soapy touch. As the layers are bonded through weak forces known as van der Waals forces, it can act as a lubricant.

Due to the presence of free electrons, it is a good conductor of electricity and heat. The melting point of graphite is 3700 °C.

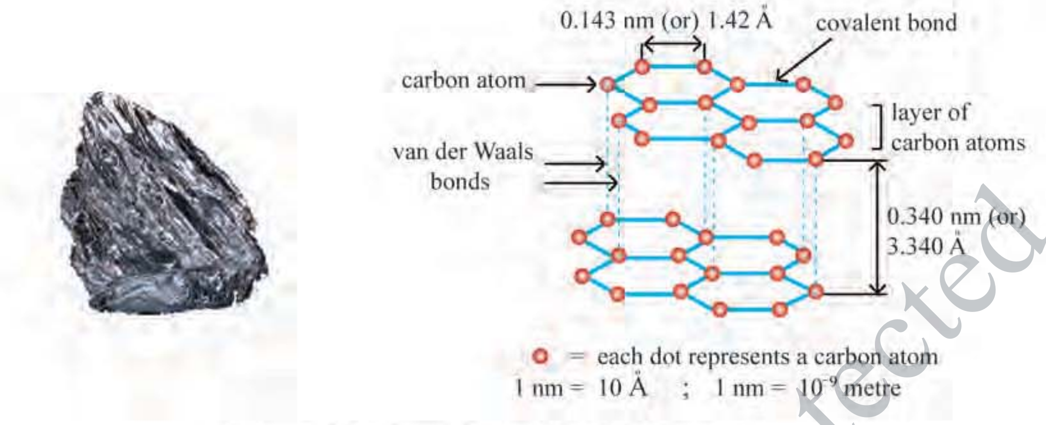


Figure 5.2 Graphite and Its Structure

Fullerene

Fullerene (C_{60}) is an allotrope of carbon in the form of a hollow sphere, ellipsoid, tube and many other shapes and sizes. Spherical fullerenes, also referred to as Bucky balls, resemble the balls used in association football. Cylindrical fullerenes are also called carbon nanotubes (Figure 5.3). Fullerenes are stable, but not totally unreactive. Fullerenes cannot be found in nature. They are the synthetic allotropes of carbon. Fullerenes have lower melting and boiling points than diamond and graphite.

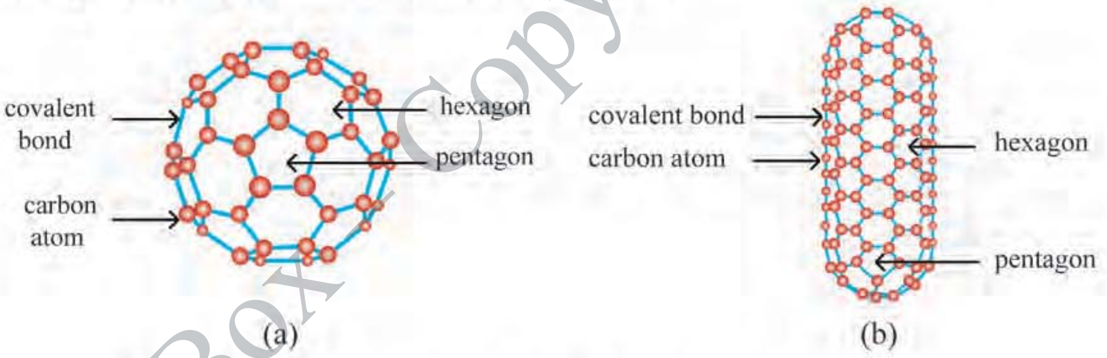


Figure 5.3 Fullerenes (a) Bucky Ball (spherical fullerene)
(b) Carbon Nanotube (cylindrical fullerene)

Graphene

Graphene is an allotrope of carbon. Graphene cannot be found in nature. It is a single layer (monolayer) of graphite (Figure 5.4). It is tightly bound in a hexagonal ring structure. Its crystalline structure is two-dimensional.

Graphene has many properties. In proportion to its thickness, it is about 100 times stronger than the strongest steel. Graphene is a transparent and flexible conductor so that it is widely used for various material/device applications, including solar cells, light-emitting diodes (LED), touch panels and smart windows or phones.

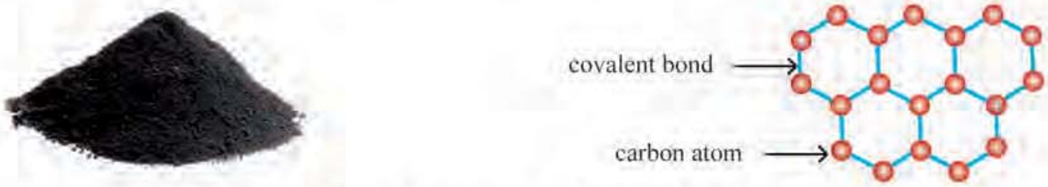


Figure 5.4 Graphene and Its Structure

(b) Other Forms of Carbon

Charcoal, coal, coke and carbon black (soot) are assumed to be amorphous forms of carbon. Now it is found that these forms of carbon contain randomly oriented small crystals of graphite.

Charcoal is made by heating wood in the absence of air. It has a porous structure and has many small holes.

Coal is found in nature. In coal the element carbon is mixed with compounds of other elements. Coal is a black heavy solid.

Coke is formed by heating coal in the absence of air. Coke is also a black heavy solid. It is almost pure carbon.

Carbon black (soot) is a black powder. When kerosene is burnt in a limited amount of air, hydrogen from kerosene combines with oxygen from the air and the carbon is left as carbon black.

(c) Chemical Properties of Carbon

(i) When carbon burns in excess air or oxygen, carbon dioxide is formed. Carbon monoxide is formed when carbon burns in a limited amount of air or oxygen.

$$\text{carbon} + \text{oxygen (excess)} \xrightarrow{\Delta} \text{carbon dioxide}$$

$$C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$$

$$\text{carbon} + \text{oxygen (limited)} \xrightarrow{\Delta} \text{carbon monoxide}$$

$$2C(s) + O_2(g) \xrightarrow{\Delta} 2CO(g)$$

(ii) Carbon can be used as a reducing agent in the extraction of some metals. When strongly heated, carbon can reduce the oxides of zinc and other metals such as CuO, PbO and Fe_2O_3 to their respective metals.

$$\text{carbon} + \text{zinc oxide} \xrightarrow{\Delta} \text{zinc} + \text{carbon monoxide}$$

$$C(s) + ZnO(s) \xrightarrow{\Delta} Zn(s) + CO(g)$$

Review Questions

(1) Diamond is very hard whereas graphite is soft. Why?
(2) Graphite is a good conductor of electricity, but diamond is not. Why?
(3) Explain why fullerene is an allotrope of carbon.
(4) Discuss the differences between graphite and graphene in their structures.

Chemistry in Society

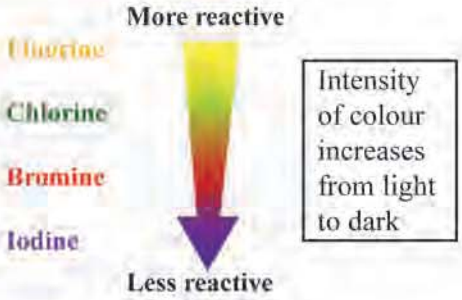
- Carbon is a major component of giant molecules called macromolecules which include proteins, lipids, nucleic acids and carbohydrates.
- There are different uses depending on the allotropes of carbon in everyday life.
- Diamond is the hardest substance known. Diamonds are used as glass cutters and drill points. Diamond is used for jewellery because of its brilliant shine.
- Graphite is used in lead pencil as it is soft. Powdered graphite is used as dry lubricant for machine parts which operate at high temperature where oil cannot be used because graphite is non-volatile. It is used in making electrode in the cells. Graphite crucibles are used as containers for melting metals at high temperature.
- Fullerene is used in artificial photosynthesis, cosmetics, surface coating of medical devices and drug delivery system.
- Graphene is widely used for solar cells, light-emitting diodes (LED), touch panels and smart windows or phones.
- Carbon (very small amount) is used to make some types of steel. Charcoal is used as a fuel for cooking. Activated charcoal is used as adsorbent in industry for bleaching (removal of colour), deodourization (removal of smell) of substances and in water purification. Coal is used as a fuel and also used to produce coke and coal tar. Coke is used as a fuel in metal industry and as reducing agent in the extraction of metals (lead, iron and zinc, etc.). Carbon black is used for making printing ink, black shoes polish and as filler in vehicle tyres and other rubber products.

Key Term

- **Allotrope** refers to two or more forms of an element that occur in the same physical state but different in properties.

5.3 HALOGENS

Halogens (F_2, Cl_2, Br_2, I_2) are diatomic molecules. Fluorine and chlorine are gases, bromine is a liquid and iodine is a solid at room temperature. They are electronegative elements. Since essential electronic structure of halogen is $ns^2\ np^5$, they are very reactive. Among them fluorine is the most reactive. Thus, none of the halogens can be found in nature in their elemental forms. They are found as salts of the halides.



(a) Preparation of Halogens

In the laboratory, halogen is prepared by heating sodium or potassium halide with manganese(IV) oxide and concentrated sulphuric acid.

Activity (2): Preparation of chlorine

Chlorine is prepared from the mixture of sodium chloride, manganese(IV) oxide and concentrated sulphuric acid on heating. It is collected by the upward displacement of air because it is slightly soluble in water and heavier than air. Due to its greenish yellow colour, it can be seen easily when the gas jar is full of chlorine gas. If required dry, the gas is passed into concentrated sulphuric acid.

$$\begin{matrix}\text{sodium}\\ \text{chloride}\end{matrix} + \begin{matrix}\text{maganese(IV)}\\ \text{oxide}\end{matrix} + \begin{matrix}\text{sulphuric}\\ \text{acid(conc.)}\end{matrix} \xrightarrow{\Delta} \begin{matrix}\text{sodium}\\ \text{hydrogen}\\ \text{sulphate}\end{matrix} + \begin{matrix}\text{manganese(II)}\\ \text{sulphate}\end{matrix} + \text{chlorine} + \text{water}$$

$$2NaCl(s) + MnO_2(s) + 3H_2SO_4(l) \xrightarrow{\Delta} 2NaHSO_4(aq) + MnSO_4(aq) + Cl_2(g) + 2H_2O(l)$$

(b) Properties of Halogens

Physical properties

Fluorine is a very pale yellow gas. It is the lightest halogen and exists as a highly toxic gas. As it is the most electronegative element, it is extremely reactive. It reacts with almost all other elements, except helium and neon.

Chlorine is a pale green gas with a choking, unpleasant smell. Chlorine is very poisonous if inhaled even in small quantities. One part of chlorine in 50,000 parts of air may be harmful. Chlorine is about 2½ times as dense as air.

Bromine is a heavy, reddish brown, volatile liquid. It has a choking, irritating smell. Bromine means 'a stench'. Liquid bromine causes burns on the flesh, which heal with difficulty. Bromine is slightly soluble in water, forming a yellowish red solution containing about 3 percent of bromine at ordinary temperature.

Iodine is a purple-black shiny solid and irritating smell. Iodine sublimes rapidly when heated, forming a violet vapour from which the black solid can again be obtained by cooling. Iodine is almost insoluble in water but readily dissolves in aqueous potassium iodide. This is due to the formation of a compound of potassium iodide and iodine, which is very soluble. This solution is brown. Iodine also dissolves in ethanol and ether, forming brown solutions, and in carbon disulphide and carbon tetrachloride, forming violet solutions.

Chemical properties

Reactivity of halogens occurs in the order of $F_2 > Cl_2 > Br_2 > I_2$.

(i) Affinity for hydrogen

Halogens react with most compounds containing hydrogen.

$$\text{hydrogen sulphide} + \text{chlorine} \longrightarrow \text{hydrogen chloride} + \text{sulphur}$$

$$H_2S(g) + Cl_2(g) \longrightarrow 2HCl(g) + S(s)$$

When a tube containing equal volumes of chlorine and hydrogen is exposed to sunlight, it explodes.

$$\text{hydrogen} + \text{chlorine} \xrightarrow{\text{sunlight}} \text{hydrogen chloride}$$
$$H_2(g) + Cl_2(g) \xrightarrow{\text{sunlight}} 2HCl(g)$$

Bromine combines with hydrogen, but not as readily as chlorine does. A mixture of bromine and hydrogen needs heat to make them combine. Also, the hydrogen bromide formed is not as stable as hydrogen chloride. Iodine has little affinity for hydrogen.

(ii) Action with metals

Chlorine reacts vigorously with metals to form metal chlorides.

$$\text{iron} + \text{chlorine} \longrightarrow \text{iron(III) chloride}$$
$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

When a very thin sheet of an alloy of copper and zinc, mainly copper, is dropped into a chlorine gas jar, it burns brightly with a green flame.

$$\text{copper} + \text{chlorine} \longrightarrow \text{copper(II) chloride}$$
$$Cu(s) + Cl_2(g) \longrightarrow CuCl_2(s)$$
$$\text{zinc} + \text{chlorine} \longrightarrow \text{zinc chloride}$$
$$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(s)$$

Bromine combines readily with most metals to form bromides. For example, copper, iron and sodium give the corresponding bromides.

Iodine is fairly active and will combine with many metals to form iodides, but it does so, much less readily than either chlorine or bromine.

(iii) Action with non-metals

Phosphorus burns spontaneously in chlorine.

$$\text{phosphorus} + \text{chlorine} \longrightarrow \text{phosphorus(III) chloride}$$
$$2P(s) + 3Cl_2(g) \longrightarrow 2PCl_3(l)$$
$$\text{phosphorus} + \text{chlorine} \longrightarrow \text{phosphorus(V) chloride}$$
$$2P(s) + 5Cl_2(g) \longrightarrow 2PCl_5(s)$$

Bromine explodes when mixed with yellow phosphorus. Phosphorus(III) bromide is made by gradually adding a solution of bromine in carbon tetrachloride to red phosphorus.

$$\text{red phosphorus} + \text{bromine} \longrightarrow \text{phosphorus(III) bromide}$$
$$2P(s) + 3Br_2(g) \longrightarrow 2PBr_3(l)$$

(iv) Bleaching action

Among halogens, chlorine and bromine have bleaching power, however, iodine does not. Chlorine reacts with water to form HOCl and HCl.

chlorine + water ⟶ hypochlorous acid + hydrochloric acid

$$Cl_2(g) + H_2O(l) \longrightarrow HOCl(aq) + HCl(aq)$$

The HOCl slowly decomposes to nascent oxygen.

hypochlorous acid ⟶ hydrochloric acid + nascent oxygen

$$2HOCl(aq) \longrightarrow 2HCl(aq) + 2[O]$$

This process is speeded up by light. The nascent oxygen bleaches dyes by oxidising them.
Chlorine will bleach moist litmus papers. This is used as a test for chlorine.
Bromine is also used as a bleaching agent but it is not as effective as chlorine.
Bromine also bleaches moist litmus papers.

(v) Oxidising properties

Halogens are oxidising agents.
Chlorine oxidises iron(II) chloride into iron(III) chloride.

iron(II) chloride + chlorine ⟶ iron(III) chloride

$$2FeCl_2(aq) + Cl_2(g) \longrightarrow 2FeCl_3(aq)$$

Bromine is also an oxidising agent, but it is not as strong an oxidising agent as chlorine. Bromine will also give majority of oxidation reactions given by chlorine.

Iodine is a mild oxidising agent, but it will not perform many of the ordinary oxidising actions attributed to chlorine and bromine. However, iodine oxidises hydrogen sulphide to form hydrogen iodide and liberate sulphur.

hydrogen sulphide + iodine ⟶ hydrogen iodide + sulphur

$$H_2S(g) + I_2(g) \longrightarrow 2HI(g) + S(s)$$

(vi) Reaction with alkalis

Halogens also react with alkalis.

Chlorine reacts on solutions of alkalis in the same way as bromine. The actions of bromine and iodine on alkalis are similar to that of chlorine.

potassium hydroxide (cold dil.) + chlorine ⟶ potassium chloride + potassium hypochlorite + water

$$2KOH(aq) + Cl_2(g) \longrightarrow KCl(aq) + KOCl(aq) + H_2O(l)$$

potassium hydroxide (hot conc.) + chlorine ⟶ potassium chloride + potassium chlorate + water

$$6KOH(aq) + 3Cl_2(g) \longrightarrow 5KCl(aq) + KClO_3(aq) + 3H_2O(l)$$

(vii) Displacement properties

When chlorine is passed into a solution of sodium bromide or sodium iodide, the respective halogen is displaced by more reactive chlorine.

chlorine + sodium bromide ⟶ sodium chloride + bromine

$Cl_2(g) + 2NaBr(aq) \longrightarrow 2NaCl(aq) + Br_2(aq)$

When bromine is passed into a solution of sodium iodide, iodine is liberated. However, iodine cannot displace chlorine and bromine from their salt solutions.

(c) Halides

A halide ion is a halogen atom with a negative charge. The halide anions are fluoride (F^-), chloride (Cl^-), bromide (Br^-) and iodide (I^-).

Test for halides

The presence of chloride, bromide or iodide ions can be tested by adding silver nitrate solution (Figure 5.5). Samples are typically acidified with dilute nitric acid to remove interfering ions, e.g. carbonate ions. Different colours of silver halides precipitates are observed.

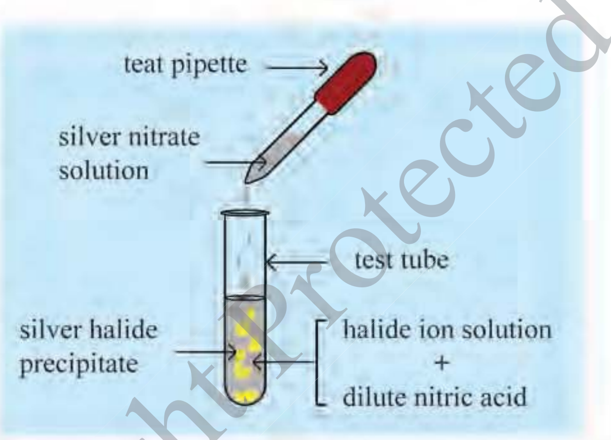


Figure 5.5 Silver Nitrate Test for Halide Ions

Chemistry in Society

- Fluorine is used in the form of fluorides in drinking water and toothpaste. It reduces tooth decay by hardening the enamel on teeth.
- Chlorine is used to make PVC (Polyvinyl chloride) plastic as well as household bleaches. It is also used to kill bacteria and viruses in drinking water.
- Bromine is used to make disinfectants, medicines and fire retardants.
- Iodine is used in medicines (e.g., to treat cases of goiter) and disinfectants (due to its antiseptic properties, e.g. 'Tincture of iodine') and also as a photographic chemical.
- Halides are used in the solder paste. It is widely used in metal halide lamps that are high-intensity discharge lamps.

Review Questions

(1) What does the term 'halogen' mean?
(2) Why are halogens highly reactive?
(3) What compound is formed when chlorine is passed over heated iron? What property does chlorine show in this reaction?
(4) Bromine reacts with sodium iodide. What property would you expect this bromine to have?
(5) What property does iodine show in the reaction with hydrogen sulphide?

Key Terms

- Siliver nitrate test is used to detect the presence of halide ions.

EXERCISES

1. Write TRUE or FALSE for each of the following statements. If FALSE, correct it.
   (a) Oxygen will burn but it cannot support combustion.
   (b) Carbon cannot exhibit allotropy or polymorphism.
   (c) Fullerene is the synthetic allotrope of carbon.
   (d) Halogens are usually found as metal halides.
   (e) Bromine can displace the chlorine from metal chloride.
   (f) Sulphur dioxide is used as fire extinguishers.
   (g) All halogens have similar properties but not identical.

2. Fill in the blanks with a suitable word or words.
   (a) When sulphur is burnt in oxygen, _____ is obtained.
   (b) Oxygen is used as a / an _____ flame in the cutting and welding of steel.
   (c) An acidic oxide is a _____ oxide.
   (d) Diamond, graphite, fullerene and graphene are the same _____.
   (e) An allotrope of carbon, _____ can be used in surface coating of medical devices.
   (f) A disease, goiter, is caused due to deficiency of _____.
   (g) In halogens, _____ is a liquid at room temperature.

3. Write equations for the following reactions that show carbon dioxide as an acidic oxide.
   (a) reaction with water (b) reaction with sodium hydroxide solution

4. Give two examples to show that carbon has reducing properties.

5. What happens when
   (a) chlorine is passed into a potassium iodide solution?
   (b) a very thin sheet of an alloy of copper and zinc is dropped into a chlorine gas jar?
   (c) chlorine is passed over heated iron?
   (d) bromine reacts with cold dilute sodium hydroxide solution?
   (e) iodine vapour is passed into hydrogen sulphide?

6. Write chemical equations in words and symbols for the following chemical reactions:
   (a) Oxidation reaction of chlorine
   (b) Displacement reaction of bromine
   (c) Affinity for hydrogen on iodine
   (d) Reaction of bromine with concentrated potassium hydroxide solution
   (e) Reaction of sodium iodide with silver nitrate solution
   (f) Reaction of iodine with dilute potassium hydroxide solution

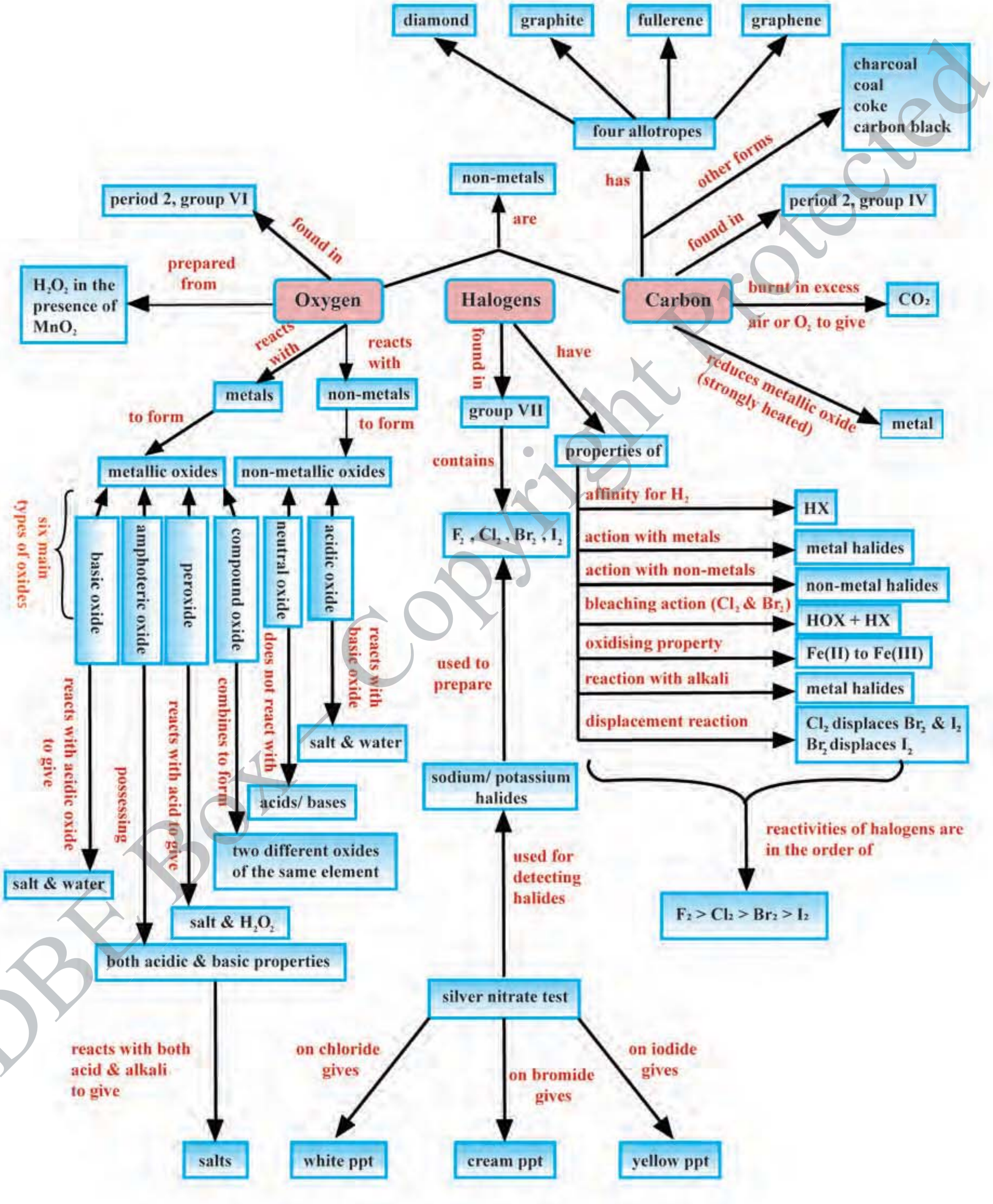
7. Write chemical equations for the preparation of bromine and iodine in laboratory.

8. How can you test the presence of chloride, bromide or iodide in a solution?

9. Halogens are strong oxidising agents. Explain with chemical equations.

10. Why is chlorine added to swimming pool water?

CHAPTER REVIEW
(Concept Map)



CHAPTER 6

ACIDS, BASES AND SALTS

Acid-Base chemistry is important in a wide variety of everyday life. In our bodies, in our home and in our industrial society, acids, bases and salts play key roles.

In our bodies, proteins, enzymes, blood, genetic materials and other components of living matter contain both acids and bases.

The organs of human and animals also contain acids. You probably know how painful a bee sting or an ant bite can be. The pain is caused by an acid called methanoic (formic) acid. The pain we sometimes feel in our leg muscles during exercise is caused by lactic acid. Our stomach produces an acid (HCl) for food digestion.

Most important mineral acids such as sulphuric acid, hydrochloric acid and nitric acid are used in industries and laboratories. In our home, many cleaners contain acids or bases. For instance, a floor cleaner acid often contains hydrochloric acid and a glass cleaner base is ammonium hydroxide. Ordinary battery acid is sulphuric acid.

H_2SO_4
(battery acid)

HCl

HNO_3

CH_3COOH
(vinegar)

Antacid
tablets

NH_4OH
(glass cleaner)

Learning Outcomes

After completing this chapter, students will be able to:

- describe the physical and chemical properties of acids, bases, alkalis and salts and their uses in daily life;
- distinguish between bases and alkalis;
- relate the role of indicators and the pH scale;
- classify the salts based on acids used and describe the preparation of salts;
- distinguish between soluble salts and insoluble salts.

We often use salts in our home. We sprinkled sodium chloride on our food to bring out its taste. We may use bath salts to help us relax in the bath and some of the medicines we take are salts. Salts are used as a preservative in pickles and in curing meat and fish, in the manufacture of soap, keeping ice from melting and making chemicals like washing soda, baking soda, etc.

Svante Arrhenius Theory (1887)

Acid is a substance which when dissolved in water produces hydrogen ions (H^+). In other words, an acid increases the number of H^+ ions in an aqueous solution.

Base is a substance which when dissolved in water produces hydroxide ions (OH^-). In other words, a base increases the concentration of OH^- ions in an aqueous solution.

6.1 ACIDS AND THEIR PROPERTIES

Many 'acids' are corrosive, meaning they destroy body tissue and clothing and many are also poisonous. Acids can be found in many foods we eat. Some organic acids are used in food preservative, food fermentation, salad, etc., such as ethanoic acid (acetic acid). Some organic acids are found in the food presented in Figure 6.1.

Yogurt (lactic acid) Apples (malic acid) Citrus fruits (citric acid) Grapes (tartaric acid) Gooseberries (Zee Phyu) (ascorbic acid) Roselle plant (Chin Baung) (oxalic acid)

Figure 6.1 Occurrence of Some Organic Acids in Nature

(a) Acids

The word 'acid' comes from the Latin word *acidus*, which means sour. Acids can be classified as mineral acids (inorganic acids) and organic acids. Acids can be strong or weak. Strong acid cannot dissociate itself without water. When a strong acid dissolves in water, it completely dissociates to produce hydrogen ions, which are protons, (H^+), and a weak acid only partially dissociates in water. Mineral acids are strong acids and organic acids are weak acids. Some acids and their dissociation reactions in water are described in Table 6.1.

Table 6.1 Names and Formulae of Some Acids and Their Dissociation Reactions in Water

Name of acids	Chemical formula	Dissociation reaction in water	Strength of acids
hydrochloric acid	HCl	$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$	strong
sulphuric acid	H_2SO_4	$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$	strong
nitric acid	HNO_3	$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$	strong
ethanoic acid	CH_3COOH	$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$	weak

$\longrightarrow$ completely dissociates in water $\rightleftharpoons$ partially dissociates in water

(**Caution**: Always add strong acid slowly to water. This is because the acid becomes very hot and splashing may happen.)

Properties

An acid is a compound which becomes a proton (H^+) donor when dissolved in water. The properties and reactions of an acid are due to these hydrogen ions.

Physical properties

(i) Acids are hazardous, irritant and corrosive.
(ii) Acids have a sour taste.
(DON'T TASTE, DON'T TOUCH.)
(iii) Acids dissolve in water to form solutions which conduct electricity.
(iv) Acid solutions have pH values less than 7.
(v) Acids have the ability to change the colour of indicators and turn blue litmus paper (an indicator) red.

Chemical properties

(i) Acid reacts with metals to form a salt and hydrogen.

$$\boxed{\text{metal} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{hydrogen}}$$

$$\text{calcium} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{hydrogen}$$
$$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$$

(ii) Acid reacts with metal oxides and hydroxides to form a salt and water only.

$$\boxed{\text{metal oxide} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{water}}$$

$$\text{calcium oxide} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{water}$$
$$CaO(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l)$$

$$\boxed{\text{metal hydroxide} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{water}}$$

$$\text{calcium hydroxide} + \text{sulphuric acid (dil.)} \longrightarrow \text{calcium sulphate} + \text{water}$$
$$Ca(OH)_2(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + 2H_2O(l)$$

(iii) Acid reacts with carbonate to form a salt, carbon dioxide and water.

$$\boxed{\text{metal carbonate} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{carbon dioxide} + \text{water}}$$

$$\text{calcium carbonate} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{carbon dioxide} + \text{water}$$
$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Chemistry in Daily Life

Some examples of most common uses of acids in daily life are listed in the following Table:

Acids	Formula	Uses
sulphuric acid	H_2SO_4	extraction of some metals such as copper, manufacture of fertilisers, detergents, paints, rubber, paper and pulp industry, car batteries and rust removal
hydrochloric acid	HCl	to help swimming pools be free of algae, to make aqua regia for dissolving gold and platinum
nitric acid	HNO_3	making fertilisers and explosives, to make aqua regia (a mixture of one part of the concentrated nitric acid and three parts of the concentrated hydrochloric acid) for dissolving gold and platinum
phosphoric acid	H_3PO_4	making fertilisers and rust inhibitor
carbonic acid	H_2CO_3	in fizzy drinks
citric acid	$C_6H_8O_7$	in fruit juices, in the preparation of effervescent salts, as a food preservative
ethanoic acid	CH_3COOH	in vinegar, used in salad dressings

Review Questions

(1) After rubbing an old copper coin with lemon juice, what visible change happens to the coin? Why?
(2) How can you detect whether a solution is acidic or not? **(Not to taste)**
(3) Ant bite is painful. Why is it so?
(4) Why can you treat bee stings with baking powder?
(5) In a laboratory, solution **A** is prepared by dissolving 10 mL of hydrochloric acid in 100 mL of water and solution **B** is prepared by dissolving 1 mL of hydrochloric acid in 100 mL of water. Which one is more concentrated? Which one is strong or weak or not?

Key Terms

- An **acid** is a compound that dissolves in water to produce hydrogen ions, H^+.
- A **dissociation reaction** is a chemical reaction in which a compound breaks apart into two or more parts.
- **Strong acid** is an acid that completely dissociates in water and gives H^+ ions. All strong acid molecules become ions in the water.

- **Weak acid** is an acid that partially dissociates in aqueous solution and gives H^+ ions. Most of these molecules remain unchanged in the water.
- An acid solution which contains the pure acid or predominantly large proportion of the acid is called **concentrated acid**.
- An acid solution which contains a relatively small amount of the acid is called **dilute acid**.

6.2 BASES, ALKALIS AND THEIR PROPERTIES

Bases and alkalis are found in many cleaning agents such as soap and many household detergents. When wood ashes are burnt, the product is alkaline. The word **alkali** comes from the Arabic 'al-qili' which means burnt ashes. It is used traditionally by gardeners as a good source of potash.

(a) Bases

A base is usually a metallic oxide or hydroxide and will react with an acid to form a salt and water only. For example,

$$\text{copper(II) oxide} + \text{sulphuric acid (dil.)} \longrightarrow \text{copper(II) sulphate} + \text{water}$$
$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

$$\text{sodium hydroxide} + \text{hydrochloric acid (dil.)} \longrightarrow \text{sodium chloride} + \text{water}$$
$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

(b) Alkalis

An alkali is a base that is soluble in water. An example of a soluble base is sodium oxide.

$$\text{sodium oxide} + \text{water} \longrightarrow \text{sodium hydroxide}$$
$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Alkalis can be strong or weak. Strong alkalis dissolve in water to produce OH^- ions in solution. Sodium hydroxide and potassium hydroxide are examples of strong alkalis. Ammonium hydroxide is the most common example of a weak alkali.

$$\text{sodium hydroxide} \longrightarrow \text{sodium ion} + \text{hydroxide ion}$$
$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

$$\text{ammonium hydroxide} \rightleftharpoons \text{ammonium ion} + \text{hydroxide ion}$$
$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Most bases are insoluble in water. MgO, CuO, Fe_2O_3, etc. are insoluble bases. They do not react with water and also not dissolve in water. Thus, it is a base and not an alkali. Some common alkalis and bases are described in Table 6.2.

Table 6.2 Some Common Alkalis and Bases

Name	Formula	Alkalis (soluble bases)	Insoluble bases
sodium oxide	Na_2O	alkali	-
sodium hydroxide	$NaOH$	alkali	-
potassium oxide	K_2O	alkali	-
potassium hydroxide	KOH	alkali	-
calcium oxide	CaO	alkali	-
calcium hydroxide	$Ca(OH)_2$	alkali	-
copper(II) oxide	CuO	-	base
magnesium oxide	MgO	-	base
iron(III) oxide	Fe_2O_3	-	base

Properties

Physical properties

(i) Strong bases are hazardous to handle.
(ii) Bases have a bitter taste and soapy feel. **(DON'T TASTE)**
(iii) Bases cause a colour change in indicators. Litmus changes from red to blue in a basic solution.
(iv) Alkalis have pH values greater than 7.

Chemical properties

(i) Bases react with acids to neutralise each other and form a salt and water. For example,

magnesium oxide + sulphuric acid ⟶ magnesium sulphate + water

$$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

sodium hydroxide + sulphuric acid ⟶ sodium sulphate + water

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

(ii) When alkalis are gently warmed with ammonium salt it gives off ammonia gas.

$$\text{sodium hydroxide} + \text{ammonium chloride} \xrightarrow{\Delta} \text{sodium chloride} + \text{water} + \text{ammonia}$$

$$NaOH(aq) + NH_4Cl(s) \xrightarrow{\Delta} NaCl(aq) + H_2O(l) + NH_3(g)$$

(iii) Alkalis react with fatty acids to form soaps.

Chemistry in Daily Life

Some common bases and alkalis and their uses are described in the following Table:

Bases and alkalis	Formula	Uses
sodium hydroxide	$NaOH$	making soap, paper, baking soda, oven cleaners
calcium hydroxide (slaked lime)	$Ca(OH)_2$	treating acidic soil (liming), making cement, limewater, mortar, plaster
calcium oxide (quicklime)	CaO	making cement
magnesium oxide	MgO	in antacids (gastric medicine), in toothpaste
ammonia	NH_3	in many household cleaners and production of fertilisers

Review Questions

(1) Oven cleaner can remove the dirt and grease from oven. What is the active ingredient in the cleaner? What is the function of that ingredient?

(2) Toothpaste contains aluminium hydroxide which removes plague. What does this tell you about the nature of plaque and bacteria on your teeth?

Key Terms

- A **base or an alkali** is a chemical compound that combines with an acid to form a salt and water. An **alkali** is a base which is soluble in water producing OH^- ions. All **alkalis** are bases but all bases are not alkalis.
- A **strong base** is a base that completely dissociates in water producing OH^- ions. All base molecules become ions in the water.
- A **weak base** is a base that partially dissociates in water producing OH^- ions. Most of the base molecules remain unchanged in the water.

6.3 INDICATORS AND THE pH SCALE

Many brightly coloured flowers, vegetables and berries make good indicators. For example, the coloured juice extracted from red cabbage is pink in acids and green in alkalis. Hydrangea flowers are interesting natural indicators. They are blue when grown in acidic soil and pink or red when grown in alkaline soil.

(a) Indicators

Indicators are dyes, or a mixture of dyes, which change colour when they are added to acids or alkalis. Some indicators can be used to determine pH because of their colour changes somewhere along the pH scale (Figure 6.2). Litmus is red in acidic solution, purple in neutral and blue in alkaline solution.

(b) The pH Scale

A measure of the acidity or alkalinity of the solution is known as pH. The pH value can be measured by pH meter (Figure 6.3). It is a much more reliable and accurate method of measuring pH than the universal indicator paper.

Substances in the body have different pH values. Acidic conditions in the stomach (pH~1.5) are needed for good digestion. Usually the body maintains the pH of blood close to 7.4. The pH scale demonstrates the strength of an acid or alkali (Figure 6.2). Solutions and their pH values are described in Table 6.3.

Figure 6.2 The pH Scale

Table 6.3 Solutions and Their pH Values

Solutions	pH value (0 to 14)
acidic	below 7
basic	above 7
neutral	equal to 7

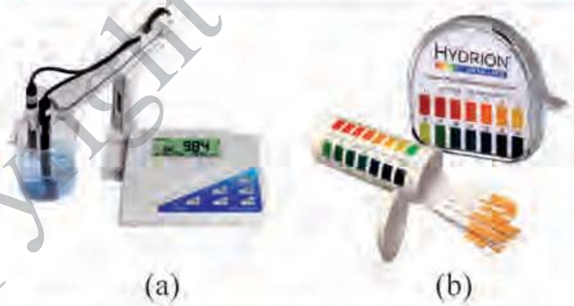

(a) (b)

Figure 6.3 (a) The pH Meter
(b) Universal Indicator Paper

Chemistry in Daily Life

- The pH is important for the correct functioning of the body, for food and water and for the growth of plants.
- Many plants do not grow properly in highly acidic or highly alkaline soil. Highly acidic soil is treated by spreading quicklime (CaO), slaked lime ($Ca(OH)_2$) or calcium carbonate ($CaCO_3$) to lower its acidity.
- Highly alkaline soil is treated by adding gypsum ($CaSO_4.2H_2O$) to lower its alkalinity.

Review Questions

(1) Which of the solutions having the following pH, are acidic or alkaline or neutral?
(a) pH 6 (b) pH 3 (c) pH 7 (d) pH 8
(2) The pH of pancreatic juice is 7.9. Is pancreatic juice acidic or basic?
(3) How do we detect whether a soil is acidic or basic?
(4) Name a common household substance with a pH (a) greater than 7 (b) less than 7 (c) almost 7.

Key Terms

- An **indicator** is a substance that has different colours in acidic and alkaline solutions.
- A measure of the acidity or alkalinity of a solution is known as its **pH**. Solutions with $pH < 7$ are acidic and those with $pH > 7$ are alkaline. The solutions of pH 7 are neutral. The pH of pure water is 7. The pH of a solution can be measured by using the pH meter.

6.4 SALTS

Many different types of salts can be found in nature. The sea water contains many salts such as sodium chloride, potassium chloride, magnesium sulphate, magnesium chloride and magnesium bromide.

The Earth's crust is made up of minerals containing various types of salts such as calcium fluoride (fluorite), magnesium sulphate (Epsom salt), lead(II) sulphide (galena) and calcium carbonate (limestone), etc.

A salt is produced when an acid reacts with a base. The salt consists of two parts. One part comes from the base, the other from the acid. An example is sodium chloride, NaCl, produced from sodium hydroxide and hydrochloric acid.

$$\text{sodium hydroxide} + \text{hydrochloric acid} \longrightarrow \text{sodium chloride} + \text{water}$$

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

The sodium ion (Na^+) part of the salt comes from the base and the chloride ion (Cl^-) comes from the acid. When an acid reacts with a base, a salt and water are formed. This reaction is known as **neutralisation**. It involves the combination of H^+ ion produced from acid and OH^- ion produced from base to form water.

$$\text{hydrogen ion} + \text{hydroxide ion} \longrightarrow \text{water}$$

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Neutralisation reaction occurs in our stomach. The acid (HCl) produced by our stomach, is so strong that it is neutralised with a base produced by cells. Salts are also produced when an acid reacts with a metal or a metal carbonate.

(a) Classification of Salts

The salts can be classified based on acids used. Some examples of salts (chloride, sulphate, nitrate, sulphite and carbonate salts) formed from different acids are shown in Table 6.4.

Some salts are soluble and some are insoluble depending on the types of metals. The examples of soluble and insoluble salts are given in Table 6.5.

Table 6.4 Some Salts Formed from Different Acids

Acids		Salts	
hydrochloric acid	HCl	**chloride salts** sodium chloride zinc chloride magnesium chloride	 NaCl $ZnCl_2$ $MgCl_2$
sulphuric acid	H_2SO_4	**sulphate salts** sodium sulphate copper(II) sulphate	 Na_2SO_4 $CuSO_4$
nitric acid	HNO_3	**nitrate salts** sodium nitrate potassium nitrate ammonium nitrate copper(II) nitrate	 $NaNO_3$ KNO_3 NH_4NO_3 $Cu(NO_3)_2$
sulphurous acid	H_2SO_3	**sulphite salts** sodium sulphite	 Na_2SO_3
carbonic acid	H_2CO_3	**carbonate salts** sodium carbonate calcium carbonate	 Na_2CO_3 $CaCO_3$
ethanoic acid	CH_3COOH	**ethanoate salt** sodium ethanoate	 CH_3COONa

Table 6.5 Soluble and Insoluble Salts

Salts	Soluble salts	Insoluble salts
nitrates	all nitrates	none
chlorides	all chlorides (except silver, mercury(I), lead(II))	silver, mercury(I), lead(II)
sulphates	all sulphates (except barium, lead(II), calcium)	barium, lead(II), calcium
carbonates	sodium, potassium, ammonium	all carbonates except those of sodium, potassium and ammonium

(b) Preparation of Salts

The preparation method depends on whether the salt is soluble in water or not. Soluble salts are usually prepared by crystallisation method. Insoluble salts are usually prepared by precipitation method.

Soluble salts

(i) Soluble salts may be prepared by using moderately dilute acids and metals. The salt formed can then be separated by crystallisation. For example,

zinc + sulphuric acid (dil.) ⟶ zinc sulphate + hydrogen

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

(ii) Salts of sodium, potassium and ammonium can be prepared from caustic soda solution (NaOH), caustic potash solution (KOH) and ammonia solution (NH_4OH), respectively, by the neutralisation using the appropriate acid.

sodium hydroxide + hydrochloric acid (dil.) ⟶ sodium chloride + water

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

potassium hydroxide + sulphuric acid (dil.) ⟶ potassium sulphate + water

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

ammonium hydroxide + nitric acid (dil.) ⟶ ammonium nitrate + water

$$NH_4OH(aq) + HNO_3(aq) \longrightarrow NH_4NO_3(aq) + H_2O(l)$$

(iii) Soluble salts can be prepared by using either the oxide or the hydroxide of the metal with the appropriate acid.

zinc oxide + sulphuric acid (dil.) ⟶ zinc sulphate + water

$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$$

zinc hydroxide + sulphuric acid (dil.) ⟶ zinc sulphate + water

$$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$$

Insoluble salts

Insoluble salts are prepared by precipitation. For example, an insoluble salt, barium sulphate, can be made by mixing solutions of barium chloride and potassium sulphate. A white precipitate of barium sulphate, $BaSO_4$, is formed.

barium chloride + potassium sulphate ⟶ barium sulphate + potassium chloride

$$BaCl_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,KCl(aq)$$

Review Questions

(1) How would you neutralise hydrochloric acid if you spill it on the floor of a laboratory?
(2) If the soil is too acidic, we add lime to the soil. Explain the purpose of this.

(3) Farmers treat the alkaline soil by using gypsum ($CaSO_4 . 2H_2O$). Why?
(4) We take gastric medicine when we feel stomach pain. Explain the action of this medicine.

Chemistry in Society

Salts play an important role in our society. Some salts and their uses in society are described in the following Table:

Salts	Formula	Uses
sodium chloride	$NaCl$	food additive
sodium sulphate sodium nitrite sodium citrate	Na_2SO_4 $NaNO_2$ $Na_3C_6H_5O_7$	food preservatives
ammonium sulphate ammonium nitrate ammonium phosphate	$(NH_4)_2SO_4$ NH_4NO_3 $(NH_4)_3PO_4$	fertilisers
potassium chloride	KCl	fertiliser
magnesium sulphate magnesium hydroxide	$MgSO_4.7H_2O$ $Mg(OH)_2$	medical uses (Epsom salt) medical uses (milk of magnesia, MOM)
calcium sulphate	$CaSO_4$	medical uses (Plaster of Paris , POP)

Key Terms

- A **salt** is a substance produced from the reaction between an acid and a base or a metal. Based on the acids, salts can be classified as chlorides, sulphates, nitrates, sulphites and carbonates, etc. Soluble salts are usually prepared from the reactions between metals and dilute acids followed by crystallisation. Insoluble salts are usually prepared by precipitation method.
- **Neutralisation** is the reaction between an acid and a base to form a salt and water only.

EXERCISES

1. Think carefully about the following statements. Are they TRUE or FALSE? If FALSE, correct it.
   (a) In general, all acid solutions contain hydrogen ions, H^+.
   (b) Copper(II) hydroxide is an alkali.
   (c) The smaller the pH value, the more acidic a solution is.
   (d) Strong acids and alkalis are harmful and corrosive.
   (e) Litmus paper can measure the range of pH of a solution.

2. Select the correct word or words given in the brackets.
   (a) Complete the following equation:
   $2KOH + H_2SO_4 \longrightarrow ?$ ($KSO_4 + H_2O$; $K_2SO_4 + H_2O$; $KSO_4 + 2H_2O$; $K_2SO_4 + 2H_2O$)
   (b) Which of the following compounds can form an aqueous solution of $pH > 7$?
   (carbon dioxide; hydrogen chloride; sodium chloride; sodium hydroxide)
   (c) Which of the following gases reacts with sulphuric acid to form a fertiliser?
   (ammonia; carbon dioxide; hydrogen; nitrogen)
   (d) A sample of pond water has a pH value of 11.
   This means that the water is (weakly acidic; neutral; weakly alkaline; strongly alkaline).
   (e) Which of the following substances could be used in excess to change the pH of soil from 5 to 7? (sodium chloride; calcium oxide; hydrochloric acid; sulphuric acid).

3. Fill in the blanks with a suitable word or phrase or numerical value with unit as necessary.
   (a) The combination of H^+ and OH^- ions to form water is called _____.
   (b) The pH of alkali solution is greater than _____.
   (c) Solutions having pH values below 4.5, turn blue litmus paper _____.
   (d) A measure of the acidity or alkalinity of a solution is known as _____.
   (e) The salt can be classified as soluble and insoluble salts depending on the types of ______.
   (f) Sodium citrate is a soluble salt. It is used as a food _____.

4. Complete the following sentences by using the words given below:
   (a) base, dissolves, hydrogen, ions, proton
   When an acid _____ in water, hydrogen _____are formed. A _____ ion is a proton. An acid is a_____ donor. It gives its proton to a _____.
   (b) hydroxides, hydrogen, dissolves, salt, oxides, water
   An acid is a compound that _____in water to produce _____ions. Acids react with metals to form _____and hydrogen. When acids react with metal _____ or _____, a salt and _____are formed.
   (c) acids, ammonia, hydroxide, salt, soluble
   Alkalis are water _____ bases. Examples of alkalis are_____ and sodium _____. Alkalis react with_____ to form a _____and water.
   (d) universal, alkaline, neutral, high, scale, seven, acidic
   The pH_____ shows how acidic or _____a solution is. Strongly_____ solutions have a low pH, strongly alkaline solution have a _____ pH. A solution that is neither acidic nor an alkaline is called a_____ solution. It has a pH of _____. The pH of a solution can be measured using _____indicator or a pH meter.

5. For each of the following pairs of solutions (of equal concentration) predict which solution has the higher concentration of H^+ ions:
   (a) HCl and CH_3COOH
   (b) H_2SO_4 and HNO_3

6. Magnesium sulphate ($MgSO_4$) is the chemical name for Epsom salt. It can be made in the laboratory by neutralising the base magnesium oxide (MgO).
   (a) Which acid should be used to make Epsom salt?
   (b) Write a balanced chemical equation for the reaction.
   (c) The acid is completely dissociated in water. Write an ionic equation. Which ion causes the solution acidic?

7. Aluminium hydroxide and calcium carbonate are often used as antacid. Write the balanced chemical equations for the reactions between these two bases and dilute hydrochloric acid.

8. Study the following diagram:

$$\left.\begin{array}{l}\text{potassium hydroxide} \xrightarrow{\text{dil. } H_2SO_4} \text{solution } \mathbf{A} \\ \text{barium hydroxide} \xrightarrow{\text{dil. HCl}} \text{solution } \mathbf{B}\end{array}\right\} \xrightarrow{\text{mixed}} \text{white ppt.}\mathbf{C} + \text{solution } \mathbf{D}$$

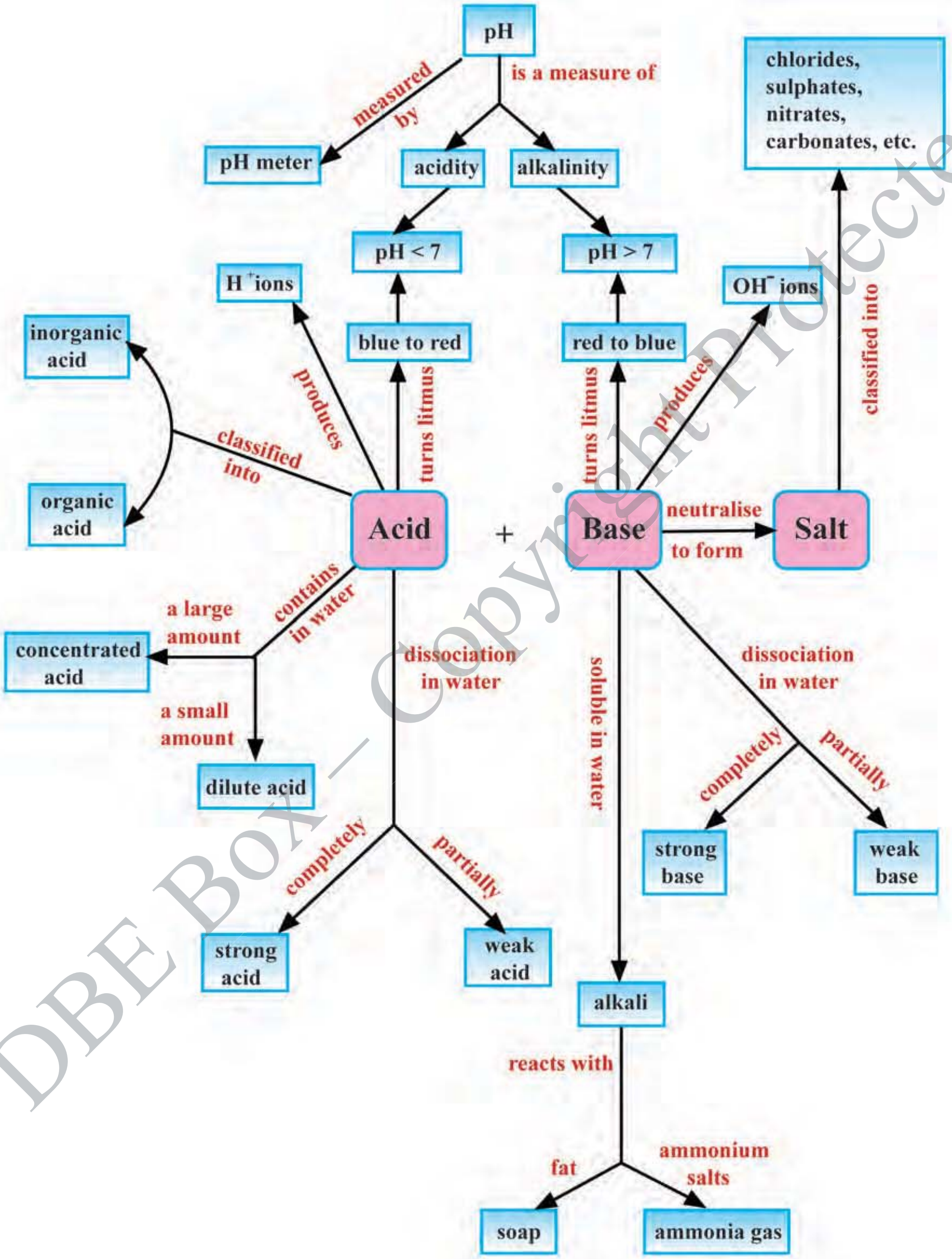
   (a) Give the names and formulae of substances **A** to **D**.
   (b) Write balanced chemical equations for the reactions taking place in the diagram.

9. You are provided three pairs of substances to produce their corresponding salts.
   (a) copper(II) oxide + dilute sulphuric acid
   (b) calcium chloride + sodium carbonate
   (c) potassium hydroxide + dilute nitric acid
   Answer the following questions.
   (i) Give the formula of each salt and predict whether the salt is soluble or insoluble.
   (ii) Which salt can be obtained by crystallisation?
   (iii) Which salt can be obtained by precipitation?
   (iv) Write a balanced chemical equation for each reaction.

10. Which two substances react to give a salt and water only? Explain.
   (a) copper(II) oxide and ethanoic acid
   (b) magnesium and dilute sulphuric acid
   (c) sodium oxide and water
   (d) zinc carbonate and dilute hydrochloric acid

11. How would you prepare the following salts? Describe their uses.
   (a) sodium sulphate
   (b) ammonium nitrate
   (c) magnesium sulphate

CHAPTER REVIEW
(Concept Map)



CHAPTER 7

AIR, WATER AND SOIL

Air, water and soil are three natural resources that we cannot live without. On the other hand, air, water and soil are the three major kinds of pollution causing harm to both living creatures and the environment. To protect our air, water and soil, one should have the knowledge related to the renewable resources such as fresh air, fresh and clean water, and fertile soil.

Learning Outcomes

After completing this chapter, students will be able to:

- recognise the air around us including the composition and the various forms of air pollution and the sources of these pollutants;
- discuss the role of various pollutants on global warming and the greenhouse effect;
- describe the Earth's surface water, both salt and fresh, including the composition, hardness and various forms of water pollution;
- explain the purification of water in terms of distillation, ion exchange and the Permutit method;
- describe the various types of soil found on the surface of the Earth, including the composition and the various forms of waste and pollutants found in the soil;
- recognise soil information including: layers, texture, composition and pH.

7.1 AIR

Without food, we could live about a month. Without water, we could live a few days. But, without air, we would die within minutes. Due to industrialisation and transportation the air is polluted with some harmful gases. Simply, we are facing today is the growing problems of acid rain, some harmful gases, global warming and ozone depletion. Polluted air is linked to a variety of health concerns, ranging from short term irritation to serious diseases. Cultural heritage sites have also suffered from enormous damage due to acid rain.

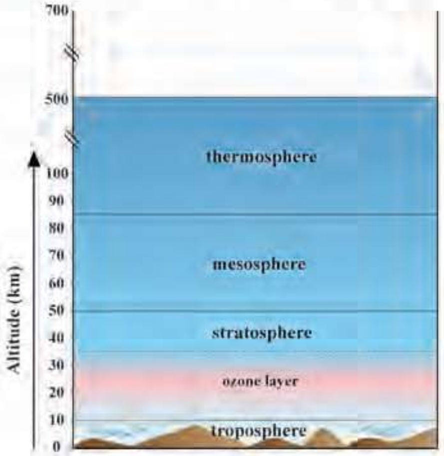


Figure 7.1 The Earth's Atmosphere

(a) The Structure of the Atmosphere

The earth is surrounded by a layer of air about 8 ~ 10 km thick called the atmosphere. The atmosphere is the blanket of gas around the Earth about 700 km. It is divided into four layers: **troposphere**, **stratosphere**, **mesosphere** and **thermosphere** (Figure 7.1). The gases in the atmosphere are held in an envelope around the Earth by its gravity. About 75 % of the mass of the atmosphere is found in the layer nearest the Earth called the **troposphere**, in which nearly all living things and nearly all human activity occur. The next region, **stratosphere**, is where we find the **ozone layer** that shields living creatures from deadly ultraviolet radiation. Beyond this layer, the atmosphere reaches into space but it becomes extremely thin beyond the **mesosphere**. The **thermosphere** is the layer in the Earth's atmosphere directly above the mesosphere.

(b) Composition of Air

Air is a mixture of several gases. The two main gases in air are nitrogen and oxygen. Other gases present in smaller amounts are carbon dioxide and the noble gases (mostly argon). As air is a mixture, its composition varies from time to time and from place to place.

Dry air contains (by volume) 78 % nitrogen (N_2), 21 % oxygen (O_2), 0.03 % carbon dioxide (CO_2) and 0.97 % noble gases. The concentration of carbon dioxide in the atmosphere has increased from 0.03 % to its present value 0.04 %. It is likely to rise as we burn more and more fossil fuels (coal, oil and gas). The amount of water vapour in air can vary widely around the world, from almost 0 % in a desert to about 5 % in a tropical forest. Figure 7.2 shows approximate composition of dry air.

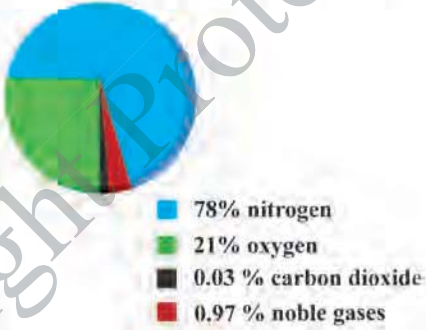


Figure 7.2 The Approximate Composition of Dry Air (by Volume)

Most of the gases in air are colourless and odourless. Some of the gases in air are essential. For example, we depend on oxygen but plants depend on carbon dioxide. And without nitrogen in air, fuels would burn too fast.

Oxygen is the reactive part of the air. It is slightly soluble in water and reacts with many other substances. The three important reactions involving oxygen are combustion, respiration and rusting. The process for respiration goes on in all our cells, by taking oxygen and releasing carbon dioxide and water.

$$\text{glucose} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water} + \text{energy}$$

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

The energy from respiration keeps us warm, allows us to move, and enables hundreds of different reactions to go on in our bodies.

There are only small amounts of carbon dioxide in the air, but it is important to all living things. Green plants need carbon dioxide for photosynthesis to produce glucose (carbohydrate) and oxygen. In this process, plants use carbon dioxide from the air and release oxygen into the air.

$$\text{carbon dioxide} + \text{water} \xrightarrow{\text{light, chlorophyll}} \text{glucose} + \text{oxygen}$$

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{light, chlorophyll}} C_6H_{12}O_6(aq) + 6O_2(g)$$

(c) Separating Gases from the Air by Fractional Distillation

Air is a mixture of gases. These gases can be separated from each other by fractional distillation. There are five steps in the separation process (Figure 7.3).

(1) Air is pumped into the plant, and filtered to remove dust particles.

(2) Carbon dioxide, water vapour and pollutants are removed since these would freeze late and block the pipes. In this step, the air is cooled until the water vapour condenses to water, followed by passing over beds of adsorbent beads to trap carbon dioxide and any pollutants in it.

(3) The air is then forced into a small space, or compressed. That makes it hot. It is cooled down again by recycling cold air, as the diagram shows.

(4) The cold, compressed air is passed through a jet, into a larger space. It expands rapidly, and this makes it very cold.

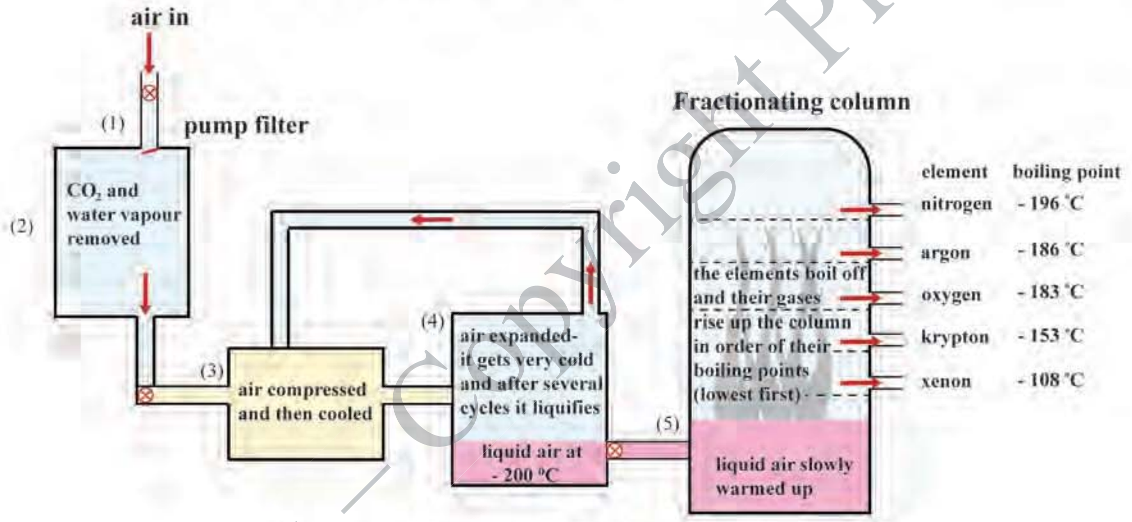


Figure 7.3 Fractional Distillation of Liquid Air

Steps 3 and 4 are repeated several times. The air gets colder each time. At -200 °C, it becomes liquid (liquefaction), except neon and helium. (These gases are separated from each other by adsorption on charcoal.)

(5) The liquid air is pumped into the fractionating column and it is slowly warmed up. The gases boil off, and are collected in tanks or cylinders. Nitrogen with the lowest boiling point boils off first.

(d) Air Pollution and Common Air Pollutants

The presence of substances in the atmosphere that are harmful to living things and to the environment contributes to air pollution. **Air pollution** is caused by solid particles (called particulates) and poisonous gases in the air. These substances are called **air pollutants**.

There are six main air pollutants: carbon monoxide, sulphur dioxide, nitrogen dioxide, methane, unburnt hydrocarbons and ozone, and their harmful effects are shown in Table 7.1.

Table 7.1 Main Air Pollutants and Their Harmful Effects

Air pollutant	Source	Harmful effects
carbon monoxide, CO a colourless gas, insoluble in water, no smell	• incomplete combustion of carbon-containing substances, e.g., charcoal, wood and petrol	• poisonous even in low concentrations • reacts with the haemoglobin in blood and prevents it from carrying oxygen around the body and will cause death
sulphur dioxide, SO_2 an acidic gas with a pungent smell	• combustion of fossil fuels in motor vehicles, power stations and factories • volcanic eruptions	• irritates the eyes and throat, and causes respiratory (breathing) problems • can form **acid rain**
nitrogen dioxide, NO_2 an acidic gas	• vehicle exhaust fumes • chemical plants • lightning activity	• causes respiratory problems • gives **acid rain**
methane,CH_4 a colourless gas, no smell	• bacterial decay of vegetable matter • cows and other farm animals when digesting food • anaerobic decomposition in natural wetlands and rice fields	• causes **global warming** because of a **greenhouse gas**
unburnt hydrocarbons	• vehicle exhaust fumes • chemical plants	• cause cancer (carcinogenic) • react with nitrogen oxides to form ozone
ozone,O_3 a colourless gas	• reaction of nitrogen oxides and unburnt hydrocarbons in the presence of sunlight	• forms **photochemical smog** which irritates the eyes and lungs and causes breathing difficulties • damages plants

(e) Global Warming and Greenhouse Effect

The Earth's surface is warmed by radiation from the sun.

Sunlight absorbed by the Earth's surface warms it up, and the surface releases heat in the form of infrared radiation. Carbon dioxide and other gases (water vapour, methane, nitrous oxide and ozone) in the air trap this radiation and prevent much of it escaping into space. The greater the amount of carbon dioxide and other heat-trapping gases, the larger is the amount of heat trapped and the warmer the Earth becomes. The average temperature of the Earth increases leading to **global warming** (Figure 7.4).

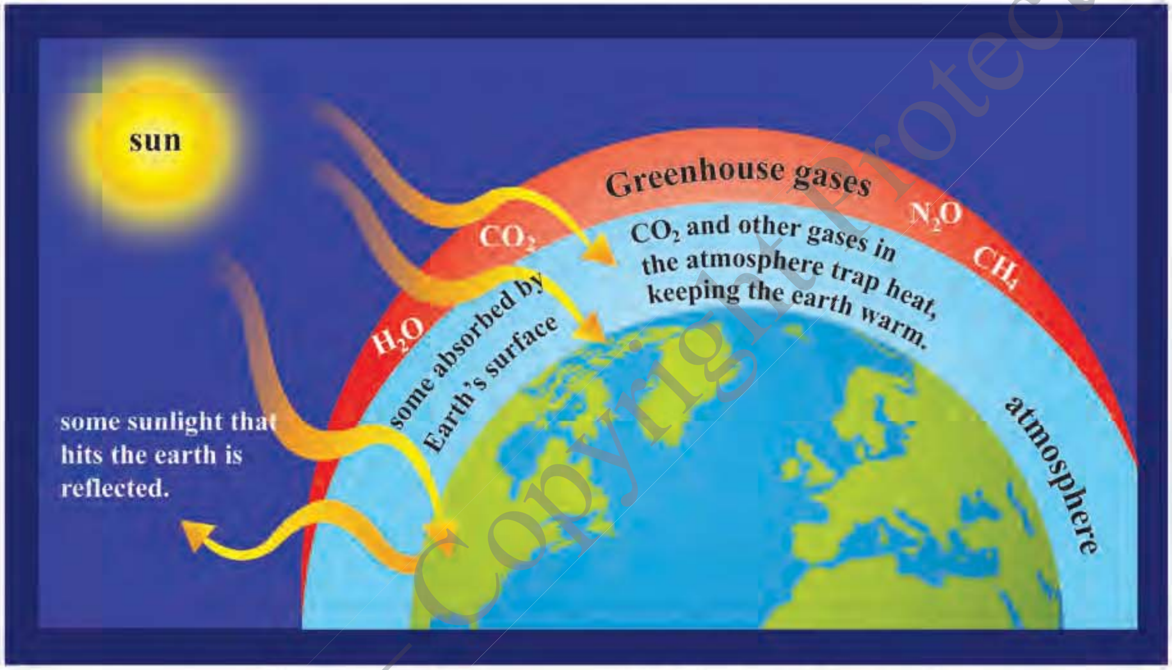


Figure 7.4 Global Warming and the Greenhouse Effect

We need greenhouse gases. Without them we would freeze to death at night, when the sun is not shining. But the level of greenhouse gases is now so high that it is causing global warming.

The gases occurred naturally in the atmosphere that trap heat are called **greenhouse gases** (GHGs) such as water vapour, carbon dioxide, methane, nitrous oxide and ozone. Besides, man-made chlorofluorocarbons (CFCs), hydrofluorocarbons (HFCs) and perfluorocarbons (PFCs), as well as sulphur hexafluoride (SF_6) are also GHGs.

Human activities, such as burning fossil fuels and farm-lands, and excessive use of fertilisers increase the amount of greenhouse gases. This greenhouse effect is gradually increasing the Earth's surface temperature, resulting in more extreme weather, such as flooding, drought, cyclone, forest fire, landslide, heat wave, etc. Another growing concern is the melting of glaciers and Artic ice which will increase sea levels resulting in many coastal communities being flooded and no longer habitable.

(f) How Acid Rain is Produced

One major environmental effect of air pollutants is the formation of **acid rain** (Figure 7.5). Rain water is naturally slightly acidic (pH of about 5.7) because carbon dioxide in the air dissolves in rain water to form carbonic acid. Sometimes, oxides of sulphur and nitrogen are released into atmosphere as industrial waste. When these dissolve in water, water becomes more acidic.

Coal-burning power plants and engines fuelled by oil or petrol release gases that can form acid rain which often falls far from its source.

$$\text{sulphur dioxide} + \text{oxygen} + \text{water} \longrightarrow \text{sulphuric acid}$$
$$2SO_2(g) + O_2(g) + 2H_2O(l) \longrightarrow 2H_2SO_4(aq)$$

$$\text{nitrogen dioxide} + \text{oxygen} + \text{water} \longrightarrow \text{nitric acid}$$
$$4NO_2(g) + O_2(g) + 2H_2O(l) \longrightarrow 4HNO_3(aq)$$

Rain water with a pH less than 5 is called acid rain.

It has many negative effects, including killing animals and plant life, and damaging metal bridges and stone buildings (Figure 7.5).

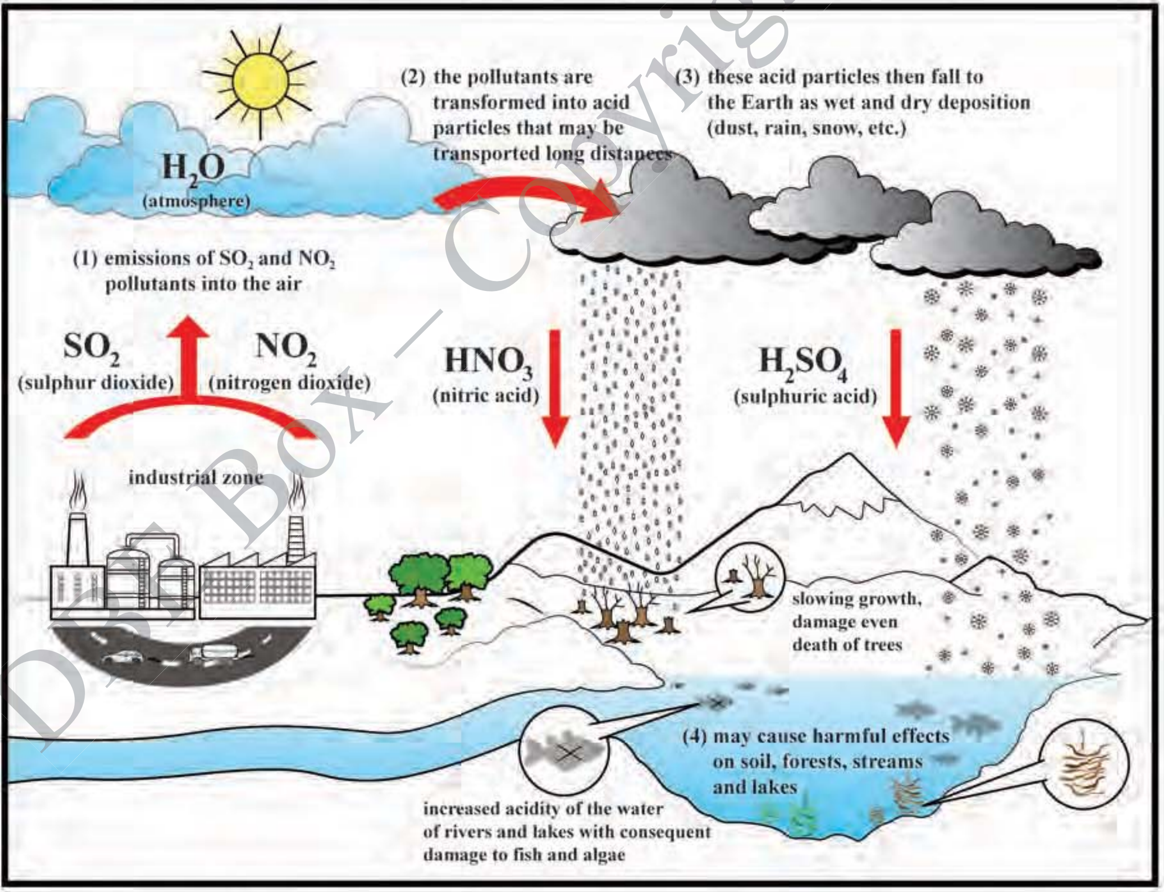


Figure 7.5 Formation of Acid Rain and its Effect on the Environment

Chemistry in Daily Life

- The useful atmospheric gases for our society include O_2, N_2, He, Ar, Ne, CO_2, etc.
- **Oxygen** is used for planes, divers, astronauts and patients with breathing problems. It is used in steel works to remove impurities (C, Si, P and Mn). Oxyacetylene flame is used as fuel for cutting and welding metals.
- **Nitrogen** is unreactive. So it is flushed through food packaging to remove oxygen and keep the food fresh. **Liquid nitrogen** (bpt. -196 °C) is used to quick-freeze food in food factories and used in hospitals to store tissue samples. Nitrogen is used with argon to fill electric bulbs because these gases do not react with tungsten filament. Nitrogen is offered as an alternative to air for tyre inflation.
- **Carbon dioxide** is important in photosynthesis. Carbon dioxide is also used widely as a coolant, a refrigerant and ingredient in the manufacture of frozen foods, and used as fire extinguisher.
- The **noble gases** are unreactive or inert. This leads to many uses.
  - Argon provides the inert atmosphere in ordinary tungsten light bulbs.
  - Neon is used in advertising signs because it glows red when a current is passed through it.
  - Helium is used to fill balloons, since it is very light, and safe.

Review Questions

(1) In the fractional distillation of liquid air, which gas is distilled over first? Why?

(2) Name two greenhouse gases. State how greenhouse gases can cause global warming.

(3) What are two pollutants that cause acid rain? Explain, using appropriate equations, how these substances are involved in the formation of acid rain.

(4) How is carbon monoxide as an air pollutant formed in the environment?

Key Terms

- **Air pollution** is the condition in which air contains a high concentration of air pollutants that may harm living things and also damage non-living things.
- The **six common air pollutants** are carbon monoxide, sulphur dioxide, nitrogen dioxide, methane, unburnt hydrocarbons and ozone, and these are harmful to health and damage the environment.
- **Global warming** is the increase in the Earth's average temperature due to the built-up of greenhouse gases in the atmosphere. Global warming may lead to melting of the polar ice caps, rise in sea levels, floods, droughts and food shortages.
- **Acid rain** is formed when acidic air pollutants mainly sulphur dioxide and nitrogen dioxide react with water in the air (atmosphere). Acid rain corrodes buildings and metal structures, damages vegetation, and kills fish in freshwater lakes and streams.

7.2 WATER

Water is the commonest compound on this planet. More than 70 % of the Earth's surface is covered with sea water, and the land masses are dotted with rivers and lakes. It is vital to our existence and survival because it is one of the main constituents in all living organisms. For example, human bones contain 31 % water, kidneys are about 82 % water and blood is about 90 % water. Those properties of water that make it uniquely suited for the support of life also make it easy to pollute. Many chemical substances are soluble in water. Removing these pollutants from our water supplies often requires enormous expenditures. You should be aware of the importance of good, safe drinking water as well as the prevention for water pollution.

(a) Occurrence of Water

Water is the most abundant substance on the Earth's surface. Out of 100 % of water that cover the Earth's surface, 97.5 % are sea water; only 2.5 % make up fresh water. Out of 2.5 % fresh water, 1.97 % make up ice caps and glaciers, 0.5 % make up ground water, only 0.02 % make up lakes and rivers, and the remaining 0.01 % are soil moisture.

Khakaborazi ice cap

Natural water in Inlay lake

Pure water cannot be found in nature. All natural waters contain impurities in varying amounts. Hence, **natural water** does not exist in a neutral state.

Rain water is the purest form of natural water. However, it may contain dissolved gases (such as oxygen and carbon dioxide) and dust from the atmosphere. Dissolved carbon dioxide makes rain water slightly acidic.

River water contains some dissolved gases and also some dissolved solids depending upon the soil over which it passes. River water is unfit to drink. The Ayeyawady River is the lifeline of Myanmar and majority of the country's population is dependent on the river for their livings. The river water finally flows into the sea carrying contaminants with it. **Sea water** contains various soluble salts (about 3.5 %).

Ayeyawady River

Sea water near Gaw Yan Gyi Island

(b) The Unique Properties of Water

(i) Pure water is a clear, transparent and colourless liquid in thin layers. Thick layers of water have bluish colour.

(ii) The freezing point of water is 0 °C (32 °F) and the boiling point of water at 1 atmosphere is 100 °C (212 °F).

(iii) Water has a greater specific heat capacity than almost any other liquid.

(iv) Water decreases in density when it freezes.

(v) With decrease in temperature, most substances diminish in volume, and hence increase in density. However, water has the very unusual property of having a temperature at which its density is a maximum. This temperature is 3.98 °C or 4 °C to the nearest degree.

(vi) Water is regarded as the most universal solvent because it dissolves almost all substances to a greater or lesser extent.

(c) Hardness of Water

The water you drink contains some dissolved solids and gases. These dissolved materials usually are not harmful and can, in fact, be good for you. Where do they come from? Rain water dissolves carbon dioxide as it falls through the atmosphere. A small fraction of this dissolved carbon dioxide reacts with the water to produce carbonic acid, which is a weak acid.

$$\text{water} + \text{carbon dioxide} \rightleftharpoons \text{carbonic acid}$$
$$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$$

As this solution passes over and through rocks containing limestone (calcium carbonate, $CaCO_3$) and dolomite (calcium magnesium carbonate, $CaMg(CO_3)_2$), the weak acid in the rain attacks these rocks, and very slowly dissolves them to form calcium and magnesium hydrogen carbonates.

$$\text{calcium carbonate} + \text{carbonic acid} \longrightarrow \text{calcium hydrogen carbonate}$$
$$CaCO_3(s) + H_2CO_3(aq) \longrightarrow Ca(HCO_3)_2(aq)$$

Some of the rocks may contain gypsum (calcium sulphate, $CaSO_4.2H_2O$), anhydrite (anhydrous $CaSO_4$) or kieserite ($MgSO_4.H_2O$), which are very sparingly soluble in water. The presence of any of these dissolved sulphates or hydrogen carbonates causes the water to become 'hard'.

(i) Soap and detergent

Soap is the sodium salt of organic fatty acid. The most common one is sodium stearate which is the sodium salt of stearic acid, $C_{17}H_{35}COOH$. The formula of sodium stearate is $C_{17}H_{35}COONa$.

sodium hydroxide + stearic acid (fatty acid) ⟶ sodium stearate + water

$$NaOH(aq) + C_{17}H_{35}COOH(s) \longrightarrow \underset{\text{soap}}{C_{17}H_{35}COONa(s)} + H_2O(l)$$

Detergents (soapless soaps) contain molecules with a salt-like group attached to a long chain of hydrocarbon. For example, sodium alkyl benzene sulphonates are synthetic soapless detergents. The structure of sodium 4-dodecyl benzene sulphonate, $C_{18}H_{29}SO_3Na$, is given below.

sodium 4-dodecyl benzene sulphonate

(ii) Effect of hard water on soap

In hard water areas all over the world, it is difficult to make the soap lather. Instead, the water becomes cloudy. This cloudiness is caused by the presence of a solid material formed by the reaction of the dissolved substances in the water with soap (basically sodium stearate) and it is a real problem. This white solid material is known as **scum**.

sodium stearate (soap) + calcium hydrogen carbonate ⟶ calcium stearate (scum) + sodium hydrogen carbonate

$$2C_{17}H_{35}COONa(s) + Ca(HCO_3)_2(aq) \longrightarrow (C_{17}H_{35}COO)_2Ca(s) + 2NaHCO_3(aq)$$

The amount of soap required to just produce a lather with water can be used to estimate the hardness of the water. To overcome the problem of scum formation, soapless detergents have been developed. The advantage of these detergents is that their salts of calcium and magnesium are soluble in water. Therefore, detergents do not form curdy or greasy scum in hard water.

sodium 4-dodecyl benzene sulphonate + calcium hydrogen carbonate ⟶ calcium 4-dodecyl benzene sulphonate + sodium hydrogen carbonate

$$2C_{18}H_{29}SO_3Na(aq) + Ca(HCO_3)_2\,(aq) \longrightarrow (C_{18}H_{29}SO_3)_2Ca(aq) + 2NaHCO_3(aq)$$

(iii) Degree of hardness

Degree of hardness of water is defined as number of parts of mass of $CaCO_3$ (calcium carbonate), equivalent to various calcium and magnesium salts present in one million parts by mass of water. It is expressed in ppm (parts per million).

Based on the degree of hardness, the water can be determined to be soft or hard. As the degree of water hardness increases, the water becomes to be hard.

Table 7.2 shows the classification of water based on degree of water hardness.

Table 7.2 Classification of Water Based on Degree of Hardness of Water

Degree of hardness of water (mg / L or ppm)	Classification
< 17.1	soft
17.1 ~ 60	slightly hard
60 ~ 120	moderately hard
120 ~ 180	hard
> 180	very hard
17.1 mg / L = 17.1 ppm	

(d) Types of Hardness

Depending upon the types of salts dissolved in water, hardness in water can be divided into two types – temporary hardness and permanent hardness.

Temporary hardness is caused by the presence of dissolved calcium or magnesium hydrogen carbonates. Temporary hardness is so called because it is easily removed by boiling.

Permanent hardness is caused by the presence of dissolved calcium or magnesium chlorides and sulphates. Permanent hardness is much more difficult to remove and certainly cannot be removed by boiling.

Effect of hardness

When water containing any of these substances is evaporated, a white solid deposit of calcium or magnesium sulphate and / or calcium carbonate (limescale) is left behind.

Calcium carbonate is formed from the decomposition of calcium hydrogen carbonate by heating. This calcium carbonate causes the '**furring**' in kettles that occurs in hard water areas. This furring may be removed by the addition of dilute acid.

$$\text{calcium hydrogen carbonate} \xrightarrow{\Delta} \text{calcium carbonate} + \text{water} + \text{carbon dioxide}$$

$$Ca(HCO_3)_2(aq) \xrightarrow{\Delta} CaCO_3(s) + H_2O(l) + CO_2(g)$$

$$\text{hydrochloric acid} + \text{calcium carbonate} \longrightarrow \text{calcium chloride} + \text{water} + \text{carbon dioxide}$$

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Blockages in hot water pipes are caused by a similar process to the furring of kettles. A thick deposit of limescale builds up.

(e) Removal of Hardness

(i) By boiling method

Temporary hardness from water is easily removed by boiling. When heated, the calcium hydrogen carbonate decomposes, producing insoluble calcium carbonate.

$$\text{calcium hydrogen carbonate} \xrightarrow{\Delta} \text{calcium carbonate} + \text{water} + \text{carbon dioxide}$$

$$Ca(HCO_3)_2(aq) \xrightarrow{\Delta} CaCO_3(s) + H_2O(l) + CO_2(g)$$

The substances in permanent hard water are not decomposed when heated and therefore cannot be removed by boiling. Both types of hardness can be removed by the following method.

(ii) By addition of washing soda ($Na_2CO_3.10H_2O$ crystals)

The calcium or magnesium ion, which actually causes the hardness, is removed as a precipitate by adding washing soda. Therefore, it can no longer cause hardness.

$$\begin{array}{ccccc} \text{calcium ion} & + & \text{carbonate ion} & \longrightarrow & \text{calcium carbonate} \\ \text{(from hard water)} & & \text{(from washing soda)} & & \\ Ca^{2+}(aq) & + & CO_3^{2-}(aq) & \longrightarrow & CaCO_3(s) \end{array}$$

(f) Water Pollution and Water Purification

The main causes of water pollution are sewage, fertilisers, pesticides, industrial wastes, oil and detergents. Pollution makes rivers and lakes smell. It kills aquatic plants and animals and other living things. It makes the water unfit for human consumption. Table 7.3 shows some sources of water pollutants and the effects.

Table 7.3 Some Sources of Water Pollutants and the Effects

Pollutants	Source	Effect
industrial waste (containing heavy metals such as Pb, Cu, Cd, Hg, Cr, etc.), dyes	industries	harmful effects (chronic health problem)
pesticides	agriculture fields	poisonous (from respiratory problem to cancer)
fertilisers	agriculture fields	waterway pollution, chemicals burn to crops, increased in air pollution, mineral depletion of the soil
oil spill	oil tanker	seriously affect the marine environment

Water pollution can be prevented by the proper disposal of sewage and industrial wastes. These pollutants should be treated and rendered harmless before they are discharged into the rivers or seas. The polluted water can be purified by several methods such as filtration, chlorination, distillation and deionisation (ion exchange), etc.

Ordinary water is more or less impure; it usually contains dissolved salts and dissolved gases and sometimes organic matter. The water is distilled away from the dissolved substances; however, it is far too expensive to be used on a large scale.

Ionic impurities can be effectively and cheaply removed from water by passing through substances like zeolite having giant structure. A zeolite is an aluminosilicate (or sodium aluminium silicate, $NaAl_2Si_4O_{12}$). It consists of a rigid framework formed by the aluminium, silicon and oxygen atoms. But sodium ions are loosely held and may be replaced by ions of calcium, magnesium and iron present in hard water. The removal of calcium, magnesium and iron results the removal of hardness according to the **Permutit method**.

$$\text{sodium zeolite} + \text{calcium chloride} \longrightarrow \text{calcium zeolite} + \text{sodium chloride}$$

$$2NaZ(s) + CaCl_2(aq) \longrightarrow CaZ_2(s) + 2NaCl(aq)$$

In order to get pure drinking water from rain water and river water, the process of water treatment used is illustrated in Figure 7.6.

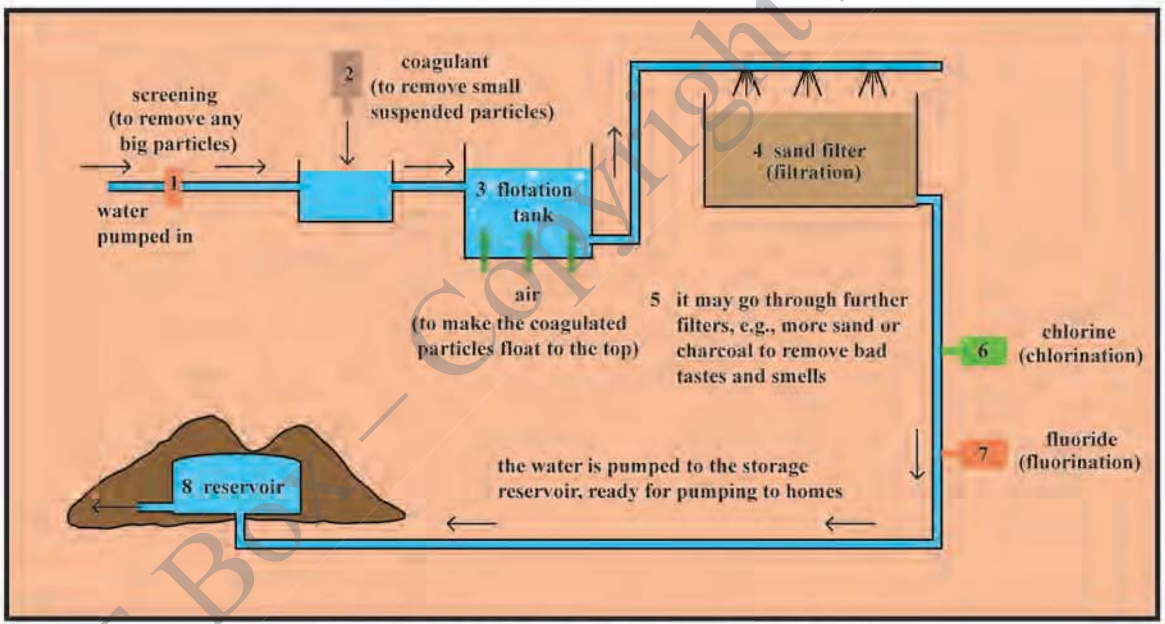


Figure 7.6 Modern Water Treatment Plant

Chemistry in Daily Life

- Water has many other important uses besides sustaining life. Water is used for cooking, cleaning, drinking, gardening and waste disposal (toilet flushing) **in home**. It is applied as a solvent as well as cleansing agent, a coolant, a major ingredient in the manufactured product and generating electricity **in industry**, and is also used for irrigation **in agriculture**.

Review Questions

(1) How will you test for hardness in water?

(2) Name two compounds in each case which can cause:
  (a) temporary hardness (b) permanent hardness.

(3) Write chemical equations, how temporary hard water reacts with soap.

(4) How do you soften temporary hard water?

(5) Explain how you will distinguish between temporary and permanent hardness in water.

(6) What are the effects of temporary and permanent hardness of water?

Key Terms

- **Hard water** is water which will not readily form a permanent lather with soap. **Soft water** is water which readily gives a permanent lather with soap.
- **Water hardness** is the amount of dissolved calcium ions, magnesium ions or both in the water. **Temporary hardness** is caused by dissolved calcium hydrogen carbonate which is removed by boiling. **Permanent hardness** is caused by dissolved calcium or magnesium chloride and sulphate which cannot be removed by boiling.
- **Degree of hardness of water** is the number of parts of mass of $CaCO_3$ (calcium carbonate), equivalent to various calcium and magnesium salts present in one million parts by mass of water (ppm).
- **Permutit method** is a process in which hard water containing calcium or magnesium salt is passed through a layer of sodium zeolite. The calcium or magnesium is removed and the corresponding sodium salt passes in solution.
- **Water pollution** is caused by the pollutants such as sewage, industrial wastes, chemical fertilisers and detergents. The treatment needed to make water fit to drink depends on the source of the water. The process of **water treatment** involves both filtration and chlorination.

7.3 SOIL

Like air and water, soil plays an essential role in our **ecosystems**. Earth's body of soil, called the **pedosphere**, has four important functions: as a medium for plant growth, as a means of water storage, supply and purification, as a modifier of Earth's atmosphere and as a habitat for organisms.

Soil is upper layer of the Earth in which plants grow, a black or dark brown material typically consisting of a mixture of organic remains, clay and rock particles. Soil plays a very important role as it produces food for human beings and animals. Good soil and a congenial climate for productivity are valuable assets for any nation. **Erosion** and a continuous cropping have taken terrific toll of the soil in many parts of the world. Excessive use of fertilisers may also cause the soil pollution. We should have the basic knowledge of soil types and general soil information since agriculture is one of the backbones of the Myanmar economy.

(a) Soil Profile

Figure 7.7 shows three main layers (top soil, subsoil, bedrock) of the soil with increasing depth. The sequence of these layers is the soil profile. Each layer has its own characteristics.

- The top layer is known as the **top soil** or the **humus layer**, which is rich in organic materials. As this layer consists of decomposed material and organic matter, it has a dark brown colour.

  The humus makes the top soil soft, porous to hold enough air and water. In this layer, the seeds germinate and roots of the plants grow. Many living organisms like earthworms, bacteria and fungi are found in this layer of soil.

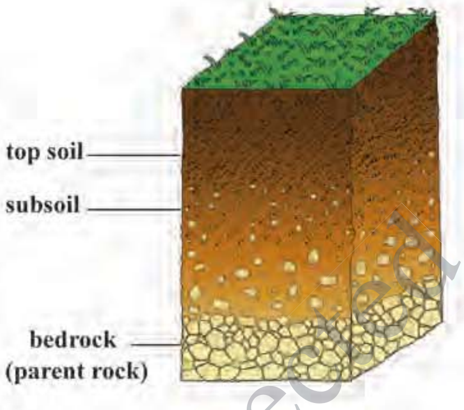


Figure 7.7 Typical Layers Found in a Soil Profile

- Just below the top soil lies another layer called **subsoil**. It is comparatively harder and more compact than top soil. It is lighter in colour than the top soil because there is less humus in this layer. This layer is less organic but is rich in minerals brought down from the top soil. It contains metal salts, especially iron oxide in a large proportion. Farmers often mix top soil and subsoil when ploughing their fields.
- The next layer is **bedrock** or **parent rock**, which lies just below the subsoil. Bedrock contains no organic matter and is made up of stones and rocks, so it is very hard.

(b) Composition of Soil

The mineral component of soil originates from the parent rocks by weathering processes, while the organic component is due to plant biomass in various types of decay as well as high populations of bacteria, fungi and animals such as earthworms.

The basic components of soil are mineral matter (45 %), organic matter (5 %), air (20 ~ 30 %) and water (20 ~ 30 %) (Figure 7.8).

Figure 7.8 Composition of Soil

Air and water occupy the pore spaces in soils. Pore spaces are the voids between the soil particles. Fine-textured soils have more total pore space than coarse-textured soils. As soils absorb water, the air space decreases. Except for gravel and rocks that occur occasionally in soils, there are three fractions; sand, silt and clay. The organic matter of soils is made up of undecomposed and partially decomposed residues of plants and animals and the tissue of living and dead microorganisms.

(c) Soil Texture

Soil texture is simply characterised by the relative proportion of sand, silt and clay separates (particles) found in the soil.

Soil texture is affected by the constituent materials found within it, specifically sand, silt and clay particles. A coarse sand will feel gritty but a wet clay will feel heavy and sticky.

Soil is made up of different-sized particles. Sand particles tend to be the biggest. Clay particles are very small. If the proportion of the sand in the soil is increased, the average size or soil particles increases and the resultant soil becomes coarser in texture.

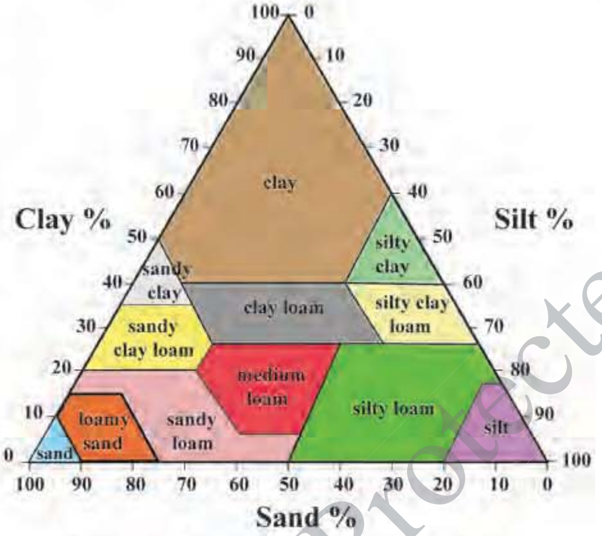


Figure 7.9 Soil Textural Triangle

If the proportion of clay in the soil is increased, the average size of the soil particle decreases and the resultant soil becomes finer in texture.

Soil textural triangle is a classification system used to determine soil classes based on their physical texture. There are twelve classes as shown in Figure 7.9.

The term **loam** refers to a **soil** with a combination of **sand**, **silt** and **clay** sized particles. For example, a **soil** with **30 % clay**, **50 % sand** and **20 % silt** is called a **sandy clay loam**.

(d) Plant Nutrients in Soil

One of the most important functions of soil in supporting plant growth is to provide essential plant nutrients – **macro-nutrients** and **micro-nutrients**. Macro-nutrients are those elements that occur in substantial levels in plant materials or in fluids in the plant. Micro-nutrients are elements that are essential only at very low levels and generally are required for the functioning of enzymes.

Essential macro-nutrients	**Essential micro-nutrients**
carbon, hydrogen, oxygen, nitrogen, phosphorus, potassium, calcium, magnesium, sulphur	boron, chlorine, copper, iron, manganese, molybdenum, zinc, sodium, vanadium

Nitrogen, phosphorus and potassium (**NPK**) are plant nutrients that are obtained from soil. They are so important for crop productivity that they are commonly added to soil as fertilisers.

Nitrogen has the most dramatic effect on the leaf growth, especially in grass and cereal plants. It is used to make protein in plants. Nitrogen bound to soil humus is especially important in maintaining soil fertility.

Nitrogen pathways in soil are shown in Figure 7.10.

Phosphorus is essential in the nucleus of every cell, so growth cannot continue in its complete absence. It is particularly associated with the development of a strong root system and floral development.

Potassium is required at high levels for growing plants. Potassium activates some enzymes and plays a key role in the water balance in plants and for some carbohydrate transformations.

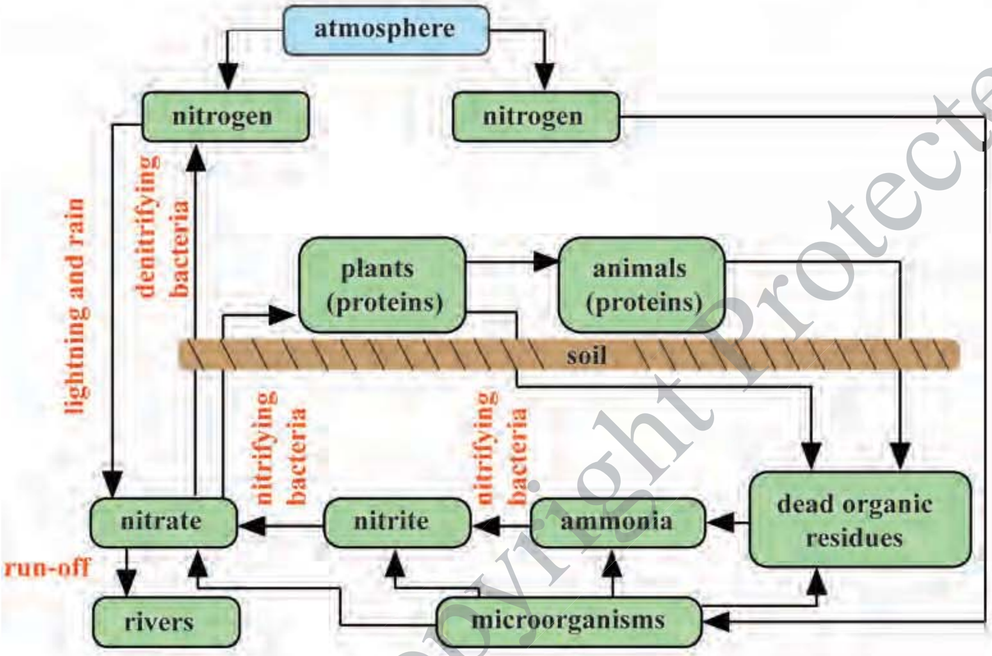


Figure 7.10 Nitrogen Pathways in Soil

(e) Soil pH

Soil pH is a fundamental property that affects a surprisingly large range of chemical, physical and biological processes in soils. Soil pH is a measure of the acidity and alkalinity in soils. pH of the soil can be measured by means of a pH meter or pH paper. The optimal pH range for most plants is **between 5.5 and 7.0**; however, many plants have adapted to thrive at pH values outside this range.

Soil pH is important because of its effect on the availability of essential elements, or nutrients, in soils. For example, many elements can be taken up by plants more easily if the soil pH is near neutral to marginally acidic. Outside this pH range, plants may be deficient in some elements, or some elements may become toxic. If the soil is more acidic than the required pH, it can be treated by adding lime. **Lime** is alkaline and it will **neutralise** the **acidity of the soil** and make it more neutral. **Alkali** or **alkaline soils** are **clay soils** with **high pH (> 8.5)**. **Gypsum** helps in the treatment of **alkaline soils**.

(f) Waste and Pollutants in Soil

Soil is the receptor of large quantities of waste products: domestic, human, animal, industrial and agricultural products.

Combustion of sulphur-containing fuels smelting process emits SO_2, and finally leaves sulphate on the soil. Nitrates from the atmosphere are deposited on the soil. Lead particulate from automobile exhausts also settles on soil along both sides of highways with heavy automobile traffic. High levels of Pb, Zn, etc. are observed on soils near lead and zinc mines, etc.

Fertilisers and pesticides applied to crops are largely retained by the soil. They become part of environmental cycles due to sorption by the soil, leaching into water, etc. Pesticides undergo degradation in soil, through the processes of biodegradation, chemical degradation, or photochemical reactions. In this respect, insects, earthworms, plants and microorganisms play important roles in biodegradation of pesticides.

Pesticide residues on crops and food products cause long-term health hazards. It may be concluded that the quality of soil has an impact on public health standards through the human food chain. The environmental health aspects of soil deserve serious attention in the near future.

(g) Sources of Soil Pollution

The main sources which pollute the soil are acid rains, repeated use or excess use of the same fertiliser, inadequate drainage system in agricultural fields, spraying the vegetable and fruit plants with insecticides and herbicides, etc.

Chemistry in Daily Life

- Soil is the foundation of basic **ecosystem** function.
- Soil filters our water, provides essential nutrients to our forest and crops, and helps regulate the Earth's temperature.

Review Questions

(1) Illustrate the typical layers found in a soil profile.
(2) What do you understand by the terms 'macro-nutrients' and 'micro-nutrients'?
(3) Why is the sandy soil not good for growing crops?
(4) By using the soil textural triangle, classify the type of soil with 10 % clay, 30 % sand and 60 % silt.

Key Terms

- **Organic matter** in soil is made up of undecomposed and partially decomposed residue of plant and animal tissues of living and dead microorganisms.
- **Soil texture** is a measure of the relative propotion of sand, silt and clay separates (particles) found in the soil.
- According to the **soil textural triangle**, the soil can be classified into twelve classes by the relative proportion of sand, silt and clay separates (particles) found in the soil.

EXERCISES

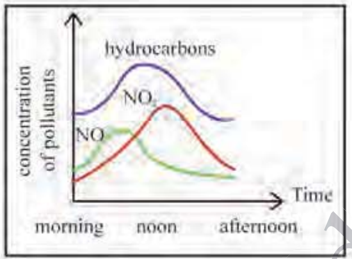
1. Write TRUE or FALSE for each of the following statements. If FALSE, correct it.
   (a) Carbon dioxide and argon are the major gases in air.
   (b) Green plants require oxygen for photosynthesis to produce glucose.
   (c) Nitrogen content in soil humus is especially important in maintaining soil fertility.
   (d) As the proportion of the sand in the soil is increased, the average size of soil particles decreases.
   (e) Most plants grow best when the soil is neutral or slightly acidic.
   (f) Nitrates from the atmosphere are deposited on the soil.

2. Tick the correct word(s), term(s), notation(s), etc., given in the brackets.
   (a) An air pollutant that can cause breathing problems is (methane; carbon monoxide; carbon dioxide; water vapour).
   (b) The approximate pH of acid rain is (less than 5; greater than 10; equal to 7; greater than 7).
   (c) What layer is called the humus layer and is made up of plant remains like leaves and twigs? (organic; top soil; subsoil; bedrock)
   (d) The important soil measurement determines how much water it can hold is (temperature; texture; colour; consistency).
   (e) Temporary hardness is easily removed by (filtration; chlorination; boiling; neutralisation).
   (f) Essential macro-nutrients for plants are (carbon; chlorine; copper; iron).

3. Fill in the blanks with suitable word(s) or phrase(s).
   (a) The two gases that cause acid rain are _____.
   (b) Nitrogen and oxygen in the air can be obtained by _____ of liquid air.
   (c) The furring of kettles or boilers is as a result of the decomposition of $Ca(HCO_3)_2$ in hard water into _____ .
   (d) Insects, earthworms, plants and microorganisms play important roles in _____ of pesticides.
   (e) Temporary hardness is so called because it is easily removed by_____ .
   (f) The layer of soil which is located at the very bottom is known as _____ .

4. Oxygen and nitrogen, the two main gases in air, are both slightly soluble in water. A sample of water was boiled, and the gases collected. The water vapour was allowed to condense and the remaining gases were measured. In a 50 cm^3 sample of these gases, 18 cm^3 were oxygen.
   (a) (i) What percentage of oxygen is present in the sample of air?
   (ii) How does this compare to the percentage of oxygen in the atmosphere?
   (b) About what percentage of atmospheric air is nitrogen?
   (c) Which gas, nitrogen or oxygen, is more soluble in water?

5. The diagram shows how the concentration of air pollutants in a busy city varies with time. How do you account for these variations with the hour of the day?



6. Discuss the following:
   (a) Industry normally requires water which has been softened.
   (b) Hard water causes kettles to fur. This 'fur' can be removed by using a dilute acid.
   (c) Hard water wastes soap.
   (d) Hard water can coat lead pipes and reduce the possibility of lead poisoning.

7. In recent years pollution of rivers and lakes has become a serious problem.
   (a) State two main sources of river pollution.
   (b) State four major water pollutants.
   (c) Suggest how to prevent these pollution.

8. Some of our drinking water is obtained by purifying river water.
   (a) Would distillation or filtration produce the purer water from river water? Give a reason for your answer.
   (b) Which process, distillation or filtration, is actually used to produce drinking water from river water?

9. Sketch the pathway of conversion of nitrogen in air to animal proteins.

10. In the treatment of water for public use, state the purpose of the addition of :
    (a) aluminium sulphate
    (b) chlorine
    (c) sodium hydroxide
    (d) sulphur dioxide

11. The soil pH is important in agriculture. Explain why.

12. Explain what is meant by the term 'pollution' with reference to air and water.

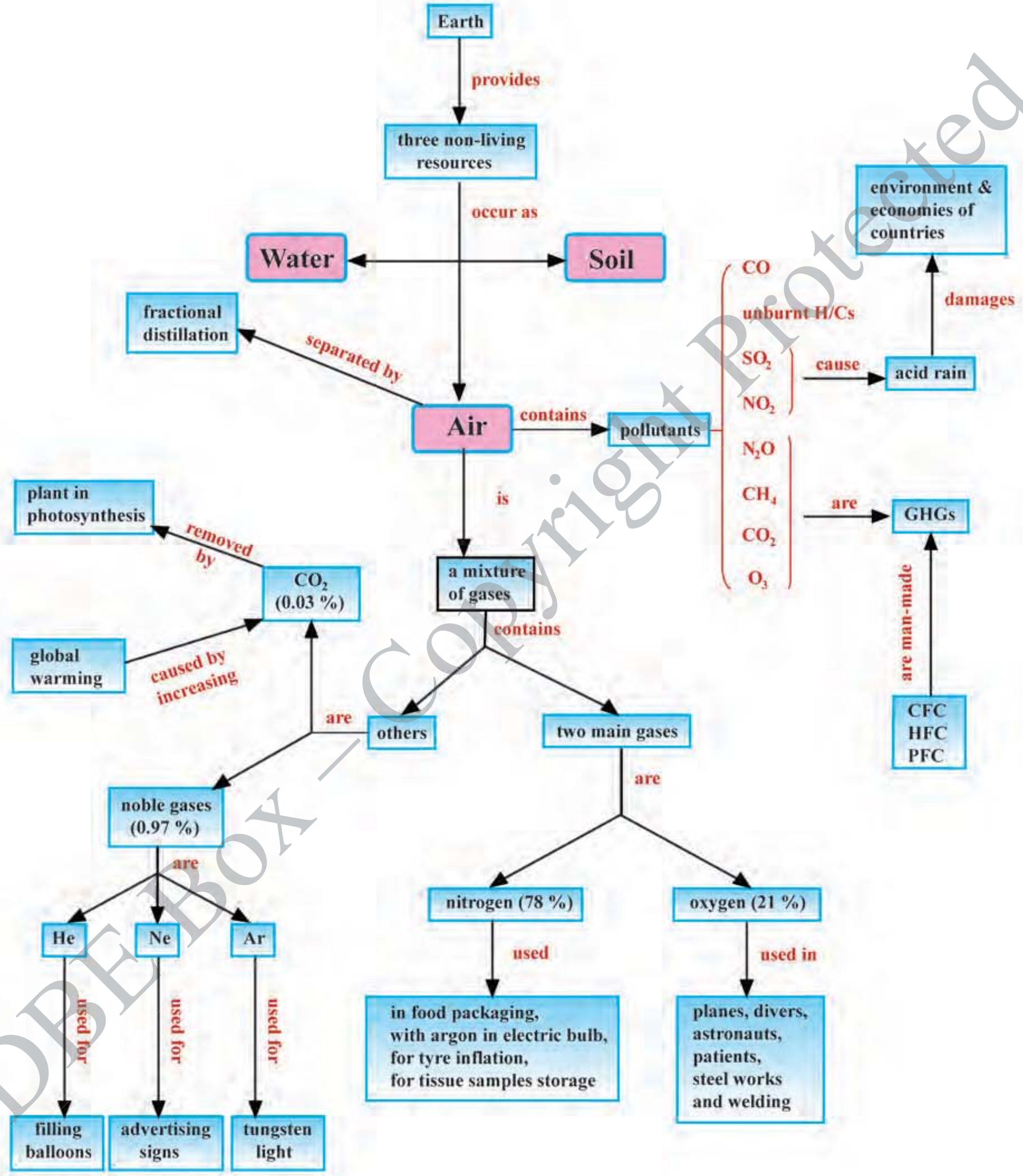
13. (a) Name an air pollutant produced by the burning of coal.
    (b) Name the air pollutant produced by the combustion of petrol in a car engine.

14. How can you treat the acidic soil? Explain briefly.

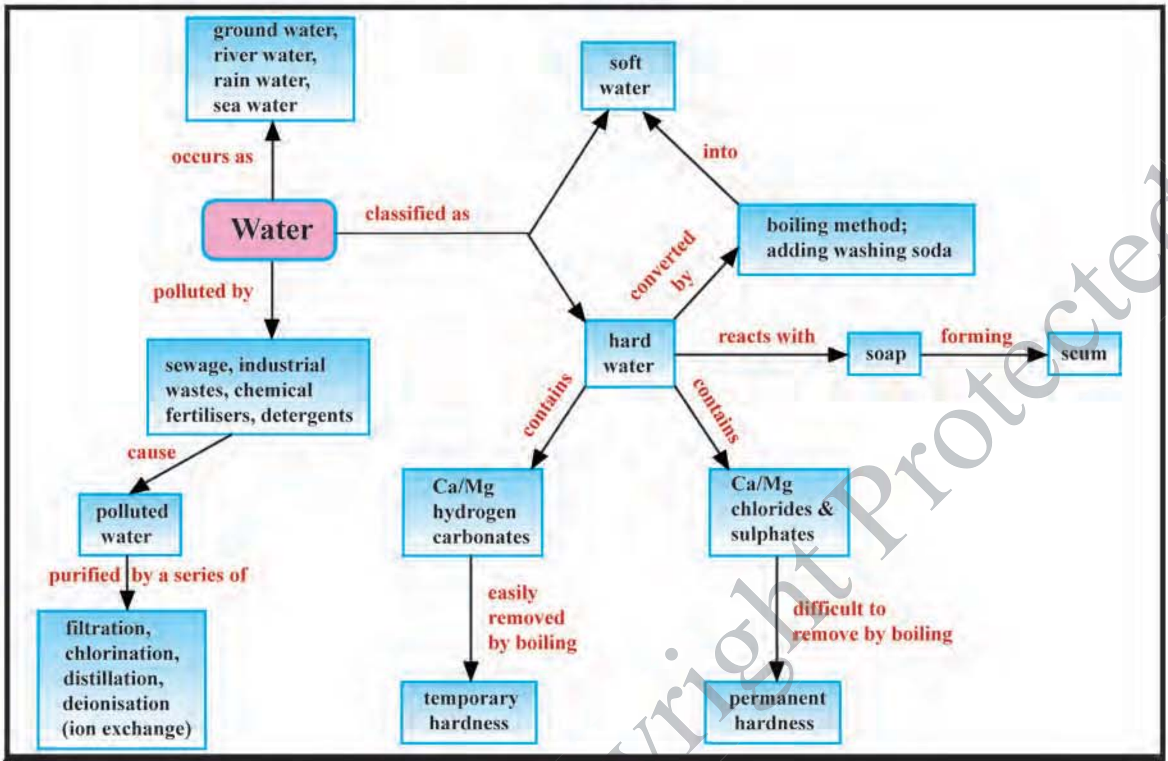
15. Make a list of four major water pollutants and explain where they come from. What damage can these pollutants do?

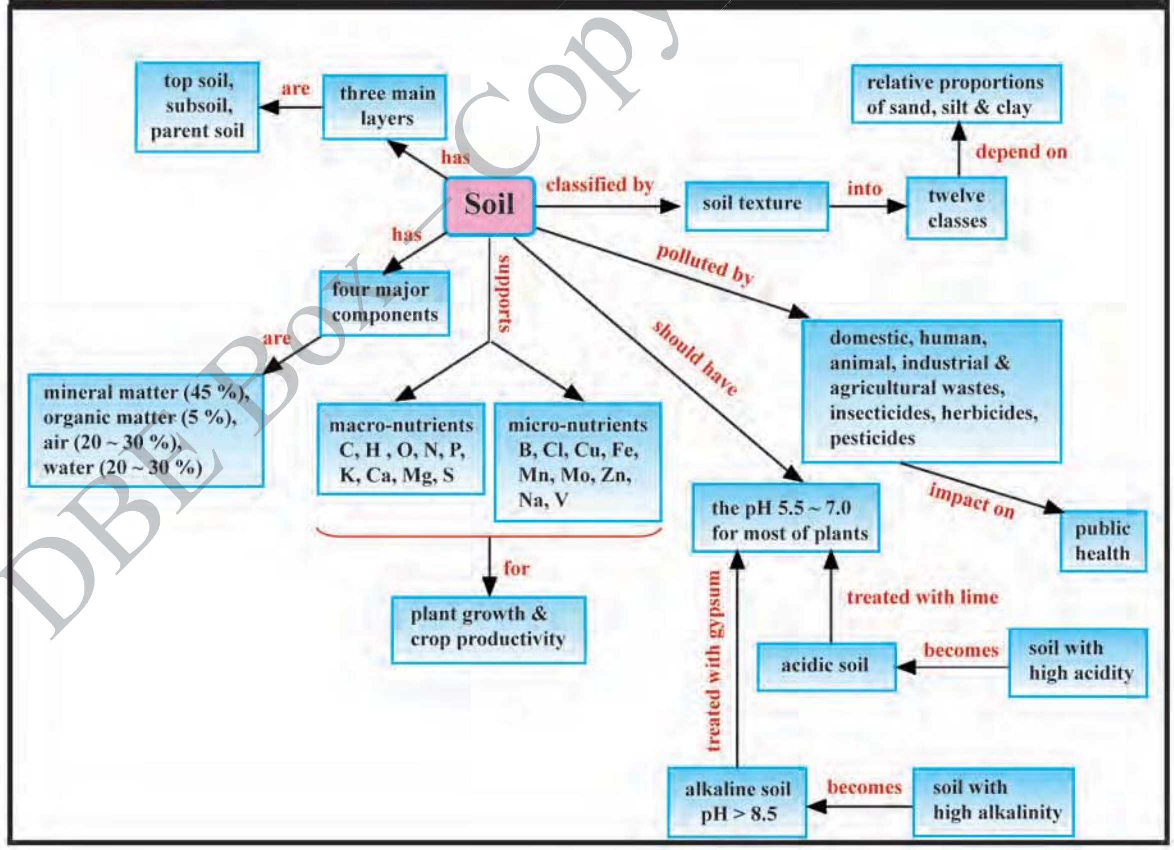
CHAPTER REVIEW

(Concept Map)



Continued from CHAPTER REVIEW (Concept Map)





CHAPTER 8

FUELS AND CRUDE OIL

In the twenty first century society, the main energy source which is used to operate the machines, power cars and buses, daily cooking our food and lighting our homes is obtained from fossil fuels.

A fuel is a substance that releases sufficient energy to do work as it undergoes a chemical change. The heat energy produced in combustion is converted into more useful forms of energy such as light energy, mechanical energy and electrical energy (electricity). Throughout history, wood and even natural gas were used thousands of years ago. They were used mainly for warmth and for cooking food. From the earlier times, wood and coal were the most common fuels. At present, fossil fuels are the main energy source used in the world.

Crude oil (petroleum), coal and natural gas are called fossil fuels because they are formed from the remains of plants and animals that lived millions of years ago. These fuels are classified as non-renewable and are finite (limited) resources because they take a very long time (millions of years) to form. Modern society is still using up fossil fuels reserved for heavy and soft industries, for non-stop transportation, generating electricity in power stations, and also for cooking.


Oil and gas production

Learning Outcomes

After completing this chapter, students will be able to:
- identify the sources, properties and behaviours of fossil fuels;
- explain the process of fractional distillation as applied to crude oil;
- describe the catalytic cracking process used to split long chain hydrocarbon molecules into shorter ones;
- recognise the sources, compositions and uses of alternative fuels, and describe the preparation methods of biodiesel and hydrogen fuel.

8.1 FOSSIL FUELS

There are three major fossil fuels. They are (a) coal (b) crude oil and (c) natural gas. Coal comes from fossil plant materials. Crude oil and natural gas are formed from the bodies of marine microorganisms. The formation of these fuels took place over many millions of years. That is why they are not only classified as non-renewable, but of finite (limited) reserved resources.

(a) Coal

Coal is fossilised plant material containing mainly carbon together with hydrogen, nitrogen and sulphur. Most coal was formed during the Carboniferous period (286-360 million years ago). The action of pressure and heat through geological forces converted the plant material in stages from peat to lignite to bituminous soft coal to hard coal (anthracite). At each stage the percentage of carbon increases. Coal contains between 80 to 90 % carbon by mass. Coal is found in many countries. The United States, Russia, China and some European countries have large coal deposits. MYANMAR also has coal deposits in Shan State, Kachin State, Taninthayi Region and Sagaing Region.

Coal is a black solid. It is mainly carbon, with small amounts of hydrogen, oxygen, nitrogen and sulphur. Coal is used in many countries to produce electricity. At a coal burning power station, coal is burnt in air to heat the water in a boiler. The steam produced turns the steam turbines to generate electricity (Figure 8.1). When coal is burnt, the main products are carbon dioxide and water. Quantities of soot, oxides of sulphur and nitrogen, and a solid residue called ash are also produced. Various kinds of pollutants are produced when coal is burnt.

Coal is also used to produce coke. When coal is strongly heated in the absence of air, a solid called **coke** is produced. Coke is almost pure carbon. It burns more cleanly than coal and it does not produce as much smoke. The main use of coke is as a reducing agent in the blast furnace for making iron.

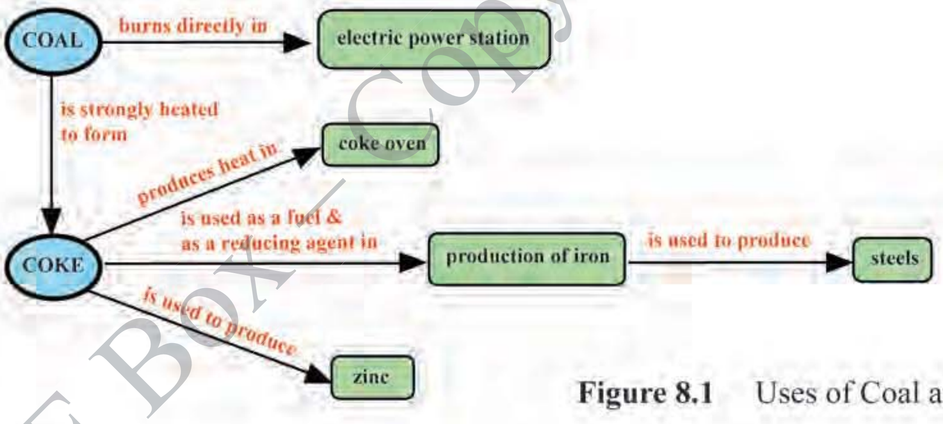


Figure 8.1 Uses of Coal and Coke

(b) Crude Oil

Crude oil (also called petroleum) is a thick black liquid. It is found together with natural gas in the Earth. Today, about 40 % of the world's energy comes from petroleum while 20 % comes from natural gas. Large amounts of petroleum are produced in the Middle East, the United States and Russia.

Myanmar is one of the world's oldest oil producers. British Burma exported its first barrel of crude oil in 1853. The London-based Burma Oil Company (BOC) was established in 1871 and began production in the Yenangyaung oil field in 1887 and the Chauk oil field in 1902.

The formation of crude oil and natural gas

The crude oil and natural gas were formed from dead animals and plants that lived in the seas a long time ago. The dead materials settled at the bottom of the sea, where it was covered with sand and other sediment. Rock then formed on top of the animal and plant remains. High pressure and temperature changed it into petroleum over millions of years. Some of it was changed into a gas - called natural gas.

Crude oil and **natural gas** are found together, held in between layers of non-porous rock in the ground (Figure 8.2). These fuels are extracted by a drilling pipe through the rock. These fuels are hydrocarbons. Hydrocarbons are made up of hydrogen and carbon only.

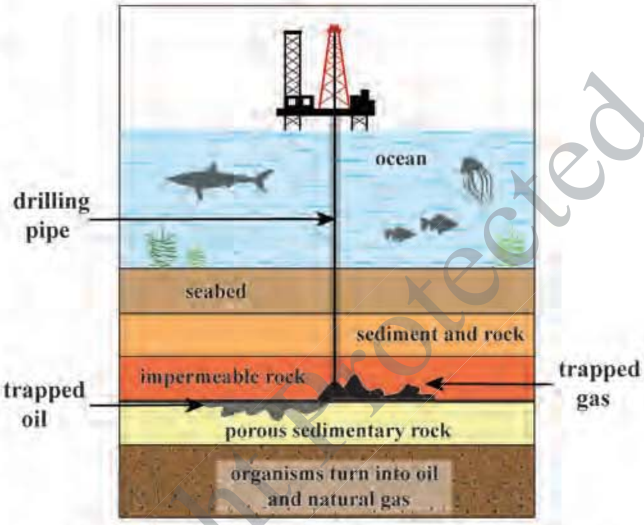


Figure 8.2 Extraction of Natural Gas and Oil from Seabed

(c) Natural Gas

Natural gas was formed at the same time as crude oil and the two are often found together, although it may occur on its own or with coal. It consists mainly of methane (85-95 %) with varying amounts of ethane, propane, butane and other gases such as carbon dioxide, nitrogen, hydrogen sulphide, etc.

Chemistry in Society

- The fuels that are derived from petroleum support more than half of the world's total energy production.
- Crude oil, coal and natural gas are non-renewable fossil fuels and contain stored energy from photosynthesis trapped millions of years ago.
- Fuel oil and natural gas are used to generate electricity. Petroleum products are used for the manufacture of synthetic fibers for clothing and in plastics, paints, chemicals, fertilisers, insecticides, soaps and synthetic rubber.
- Myanmar is today primarily a natural gas producer. As of 2015, Myanmar exports gas to Thailand and China.

Review Questions

(1) Why is petroleum called a fossil fuel?
(2) Describe the uses of coal.

Key Terms

- A **fuel** is a substance that releases sufficient energy to do work as it undergoes a chemical change.
- **Fossil fuels** consist of coal, petroleum and natural gas. **Natural gas** consists mainly of methane, CH_4. **Crude oil** (petroleum) is a mixture of many different hydrocarbon molecules.
- **Non-renewable fuels** are fuels which take millions of years to form and which are used up at a rapid rate.
- **Renewable fuels** are fuels produced from renewable resources. (e.g., vegetable oils, animal oils, etc.)
- **Combustion of coal** releases nitrogen oxides, sulphur dioxide, particulate matter (PM), mercury and dozens of other substances known to be hazardous to human health.

8.2 FRACTIONAL DISTILLATION OF CRUDE OIL

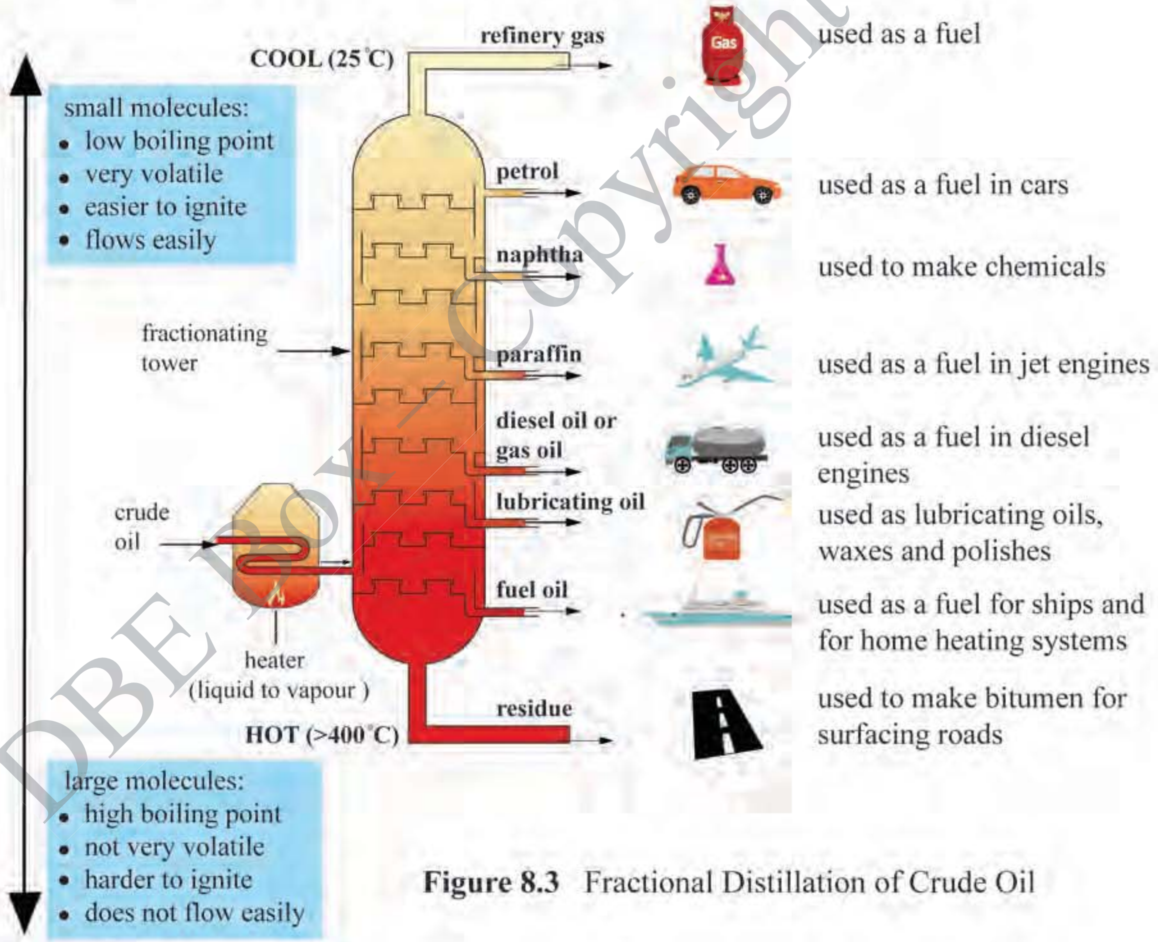


Figure 8.3 Fractional Distillation of Crude Oil

Crude oil is a mixture of many different **hydrocarbon** molecules. These molecules have different sizes and number of carbon atoms. The small molecules have few carbon atoms and low boiling points, while the large molecules have many carbon atoms and high boiling points. Therefore, it is necessary to refine the crude oil into useful fuels and chemicals.

Separation of the crude oil takes place in a fractional distillation column, or fractionating tower into different fractions (parts) in an oil refinery (Figure 8.3).

Crude oil is heated in a furnace. Many fractions could be collected, each having a different boiling point range. The oil vaporises and passes up the fractionating column. The fractions condense and come out of the column at different heights depending on their boiling points. The petroleum gas fraction comes out first at the top of the column as its molecules have the lowest boiling points. Then, a series of fractions such as petrol, naphtha, kerosene and diesel comes out in order of increasing boiling points, number of carbon atoms and viscosity. The lubricating oil fraction comes out at the bottom because its molecules have higher boiling points, followed by fuel oil. Bitumen is the residue at the bottom of the column.

All the fractions are insoluble in water and burn in air. The properties and uses of some of the main fractions from the distillation of crude oil are given in Table 8.1.

Table 8.1 Some Important Crude Oil Fractions

Fraction	**Approximate boiling point range / °C**	**Approximate number of carbon atoms per molecule**		**Important uses**
refinery gas (petroleum gases)	below room temperature < 40	1 ~ 4	**Increasing boiling point and viscosity** ↓	bottled gas for gas cookers and motor cars
petrol (gasoline)	35 ~ 75	5 ~ 10		petrol for motor cars
naphtha	70 ~ 170	8 ~ 12		petrochemicals
paraffin (kerosene)	170 ~ 250	10 ~ 14		fuel for jet aircraft; kerosene lamps for light and kerosene stoves for cooking
diesel oil	250 ~ 340	15 ~ 25		fuel for diesel engines of buses, lorries, trucks, steamers and trains
lubricating oil	350 ~ 500	19 ~ 35		lubricant in engines to reduce friction; also for making waxes and polishes
fuel oil	500 ~ 600	30 ~ 40		fuel for ships, factories and central heating
bitumen (residue)	> 600	> 70		a black substance used to make surface roads and roofing

Note: 'Crude oil' (UK) is the same as 'petroleum' (USA); 'petrol' (UK) is the same as 'gasoline' (USA); and 'paraffin' (UK) is the same as 'kerosene' (USA).

There is a greater demand for petrol and kerosene than other fractions. Consequently, cracking method is used to produce smaller molecules from larger hydrocarbon molecules.

Chemistry in Society

- Gasoline, kerosene and diesel oil provide fuel for automobiles, tractors, trucks, aircraft and ships.
- Lubricating oil is used as lubricant in engines to reduce friction; also for making waxes and polishes.
- Bitumen is used for surfacing roads.

Review Questions

(1) Name a crude oil fraction that: (a) is used for jet aircraft (b) has the smallest molecules (c) is the most viscous (d) has molecules with 19-35 carbon atoms.

(2) Consider the following petroleum fractions:
naphtha, paraffin, bitumen, diesel oil, lubricating oil
Which of the above fractions:
(a) has the lowest boiling point; (b) has the highest boiling point;
(c) is used to make waxes; (d) is used as a fuel for jet engines;
(e) contains 15 ~ 25 carbon atoms per molecule?

(3) In an oil refinery, the mixture of hydrocarbons in petroleum is separated into fractions. Petroleum fraction **X** has the boiling point range of 35-70 °C. Petroleum fraction **Y** has the boiling point range of 170-250 °C. List three facts in which petroleum fraction **X** differs from petroleum fraction **Y**. Give reasons for your answer.

(4) There is a limited quantity of petroleum on Earth. Describe two ways of conserving petroleum.

Key Terms

- **Hydrocarbon** is any of a class of organic chemical compounds composed only of the elements carbon (C) and hydrogen (H).

8.3 CATALYTIC CRACKING

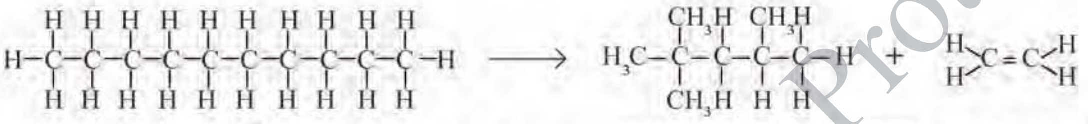
Fuels made from oil mixtures contain large hydrocarbon molecules and are not efficient. They do not flow easily and are difficult to ignite. Crude oil often contains too many large hydrocarbon molecules and not enough small hydrocarbon molecules to meet demand. Consequently, cracking is important to convert the larger hydrocarbon molecules to smaller ones.

Larger molecules from the heavier fractions (paraffin (kerosene) and diesel) can be broken into smaller, more valuable, molecules. When a catalyst is used, this process is called **catalytic cracking** ('cat cracking').

Cracking is a **thermal** decomposition process in which large alkane hydrocarbon molecules are broken down by passing them over a heated catalyst under pressure. The products are smaller alkanes used for fuels (e.g., petrol or diesel) and alkenes which are used to make polymers: plastics and other important compounds.

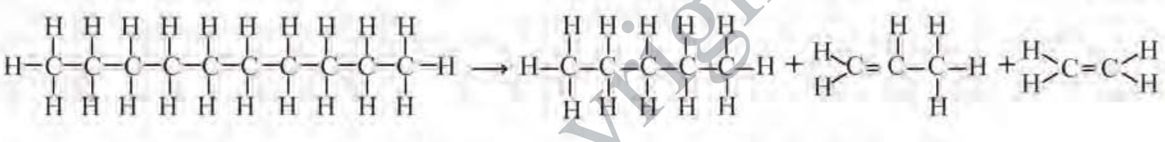
Cracking takes place in huge reactor. In this reactor, particles of catalyst (made of powdered minerals such as silica, alumina and zeolites) are mixed with the hydrocarbon fraction at a temperature around 500 °C and moderately low pressure. The cracked vapours containing smaller molecules are produced by the following types of reactions:

$$\underset{C_{10}H_{22}}{\text{Decane}} \xrightarrow[\text{moderately low pressure}]{\text{heat, catalyst}} \underset{C_8H_{18}}{\text{2, 2, 4-trimethyl pentane}} + \underset{C_2H_4}{\text{ethene}}$$



$$\text{Decane} \longrightarrow \text{pentane} + \text{propene} + \text{ethene}$$

$$C_{10}H_{22} \longrightarrow C_5H_{12} + C_3H_6 + C_2H_4$$



Cracking reactions generally give two main types of products:

(i) an alkane with a shorter chain than the original and

(ii) a short-chain alkene molecule.

Both these products are useful. The shortened alkanes can be blended with the gasoline fraction to enrich the petrol. The alkenes are useful as raw materials for making several important products. Figure 8.4 shows the various uses for the ethene produced; preparation of ethanol and plastics such as polyethene, polychloroethene and polystyrene. Here in, ethene polymerises to polyethene, i.e., many ethene molecules combine to form larger molecule polyethene that contains repeating structural units. Propene polymerises to polypropene (trade name 'polypropylene'), while butene polymerises to produce synthetic rubber.

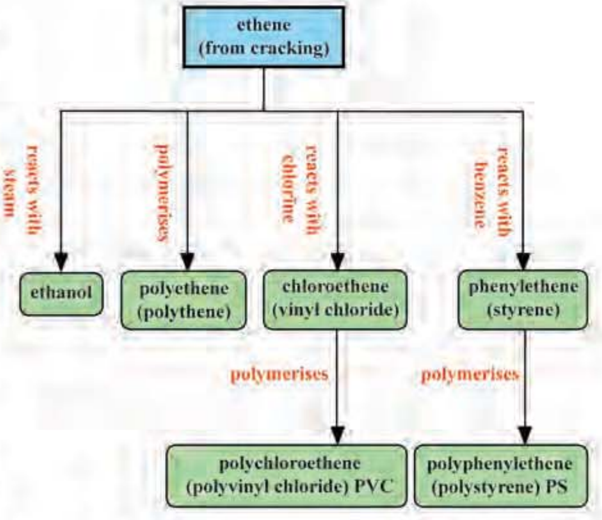


Figure 8.4 Important Products from Ethene

The alkane produced from cracking may be straight chain alkanes which turn to branched isomer on heating with the catalyst. Therefore, they have higher octane ratings (octane number) than unbranched alkanes. The higher the **octane number**, the greater the fuel's resistance to knocking in an internal combustion engine.

The octane number of a given fuel is determined by comparing the amount of knocking that fuel causes when combusted with the amount of knocking caused by two standard reference fuels; *iso*-octane which resists knocking **(antiknocking)** and has an octane number of 100 and heptane which causes knocking and has an octane number of 0. For example, if a gasoline sample has the same antiknock quality as that of a mixture containing 90 % *iso*-octane and 10 % heptane, then the octane number for that sample is defined as 90.

Chemistry in Society

- The short chain alkenes such as ethene are always produced in cracking. They are important for use in chemical industry and in the production of plastics.
- The branched-chain alkanes produced by cracking are useful components of high octane petrol.
- Antiknocking agent (e.g., tetraethyl lead) is a gasoline additive used to reduce engine knocking and increase the fuels octane rating by raising the temperature and pressure at which auto-ignition occurs.

Review Questions

(1) What happens during cracking? Discuss briefly.
(2) Cracking is a thermal decomposition. Explain why.
(3) Describe the usual conditions needed for cracking a hydrocarbon in the petroleum refinery.
(4) Explain why cracking is so important.

Key Terms

- **Catalytic cracking** is a process used to split long chain alkanes into shorter alkanes and alkenes in the presence of catalyst under pressure and high temperature.
- **Catalyst** is a substance that speeds up the chemical reaction without getting consumed.
- **Zeolites** are microporous, aluminosilicate minerals commonly used as commercial adsorbents and catalysts.
- **Thermal decomposition** is the breaking down of a chemical compound, caused by heat.
- **Octane rating (octane number)** is a measure of a fuel's ability to resist 'knock'.
- **Engine knocking** refers to the sharp sounds caused by combustion of some of the compressed air-fuel mixture in the cylinder.

8.4 ALTERNATIVE FUELS

Fossil fuels take millions of years to form. There are limited amounts in the Earth. At the present rate of consumption, petroleum and natural gas may run out within 50 years and coal will only last for a further 250 years. Therefore, scientists have tried to overcome the problem of limited crude oil supply by looking for alternative fuels to replace crude oil.

An **alternative fuel** is an internal combustion engine fuel other than gasoline or diesel oil. Alternative fuels include natural gas (methane, compressed natural gas - CNG), propane (liquefied petroleum gas - LPG), hydrogen fuel, biomass-derived fuels, biodiesel, bio-alcohols (including ethanol and methanol), alcohol mixtures with gasoline or other fuels (gasohol) and electricity.

Hydrogen fuel, biomass-derived fuels, biodiesel, bio-alcohols (including ethanol and methanol) are renewable fuels, and also known as alternative transport fuels.

Some alternative fuels and their uses are described in Table 8.2.

Table 8.2 Some Alternative Fuels and Their Uses

Fuels	Source	Composition	Uses
LPG (Liquefied Petroleum Gas)	petroleum gas	propane and butane	used as fuels in vehicles, cars, trucks and stationary power generation, for cooking and other heating systems
CNG (Compressed Natural Gas)	natural gas	90 % methane	used as fuels in vehicles, cars, trucks and stationary power generation
biodiesel	plant oils, animal oils	long chain esters	used in power tractor engines, petro-diesel engines and electricity generation engines
biogas	waste organic matter	methane	used for heating and cooking, and the solid residue is used as a fertiliser
hydrogen fuel	water, petrol and natural gas	hydrogen	used as fuels for cars, in space shuttles and other big rockets
gasohol	petrol and ethanol	90 % petrol + 10 % ethanol, 15 % petrol + 85 % ethanol (US)	used as fuels in vehicles

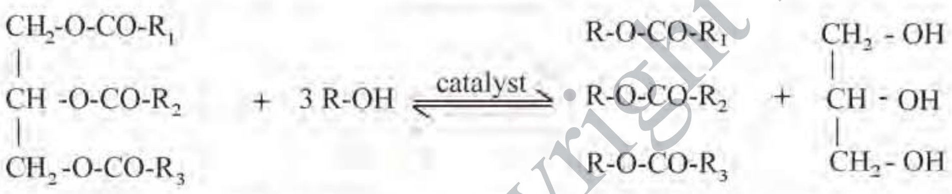
Preparation of biodiesel

The plant or animal oils have to be converted to biodiesel by **'Transesterification'**. The conversion involves four stages :

(1) the preparation of nearly **100 % pure methanol or ethanol**,
(2) the addition of **potassium hydroxide or sodium hydroxide basic catalyst** to the prepared pure methanol or ethanol,
(3) the treatment of the seed oil with the prepared basic catalytic solution and the solution heated to 60 °C which is the **transesterification** process producing methyl or ethyl ester as **the product (biodiesel)** of the reaction and
(4) the removal of glycerine and sodium or potassium salt of fatty acids (soap) from the reaction mixture by washing with water, and pure biodiesel is separated out by using the **biodiesel processor**.

The transesterification reaction can be generally presented as below.

$$\text{triglyceride} + \text{alcohol} \xrightleftharpoons{\text{catalyst}} \text{esters} + \text{glycerol}$$



$$\begin{array}{l} CH_2\text{-O-CO-}R_1 \\ | \\ CH\ \text{-O-CO-}R_2 \\ | \\ CH_2\text{-O-CO-}R_3 \end{array} + 3\ R\text{-OH} \xrightleftharpoons{\text{catalyst}} \begin{array}{l} R\text{-O-CO-}R_1 \\ \\ R\text{-O-CO-}R_2 \\ \\ R\text{-O-CO-}R_3 \end{array} + \begin{array}{l} CH_2\text{ - OH} \\ | \\ CH\text{ - OH} \\ | \\ CH_2\text{- OH} \end{array}$$

Biodiesel can be used to power tractor engines, petro-diesel engines and electricity generation engines. It is the potential substitute for petro-diesel since the source of the biodiesel is renewable and cost effective.

Hydrogen fuel

Most hydrogen is manufactured on a **large scale** in industry from petrol and natural gas. For example, a mixture of methane (from natural gas) and steam is passed over a nickel catalyst.

$$\text{methane} + \text{steam} \xrightarrow{Ni} \text{carbon monoxide} + \text{hydrogen}$$

$$CH_4(g) + H_2O(g) \xrightarrow{Ni} CO(g) + 3H_2(g)$$

The carbon monoxide gas is then reacted with more steam.

$$\text{carbon monoxide} + \text{steam} \longrightarrow \text{carbon dioxide} + \text{hydrogen}$$

$$CO(g) + H_2O(g) \longrightarrow CO_2(g) + H_2(g)$$

The hydrogen is separated from the carbon dioxide by passing the gases through an alkali to absorb the acidic carbon dioxide.

Smaller quantities of hydrogen are produced by the electrolysis of water containing sulphuric acid, though it is more expensive.

$$\text{water} \longrightarrow \text{hydrogen} + \text{oxygen}$$

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$

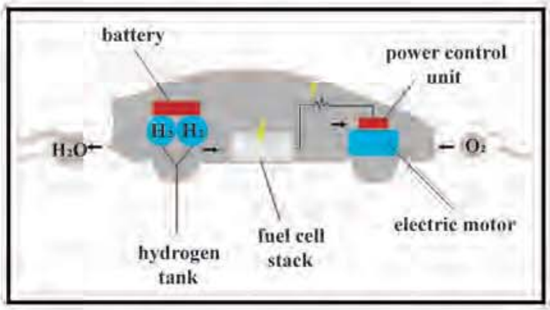
Hydrogen burns cleanly in air. The product is steam, which is a non-pollutant.

$$\text{hydrogen} + \text{oxygen} \longrightarrow \text{steam}$$

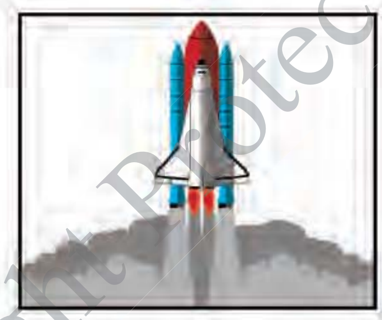
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

However, hydrocarbon fuels, such as petrol and diesel, produce polluting oxides of carbon in combustion.

Hydrogen produces at least twice as much heat energy per gram when burnt, than any other common fuel. This is why it is used as a fuel in space shuttles and other big rockets. Hydrogen has great possibilities as a fuel for cars, replacing petrol. Experimental hydrogen-powered cars are already being used.



Hydrogen fuel cell car



Hydrogen fuel cell in space shuttle

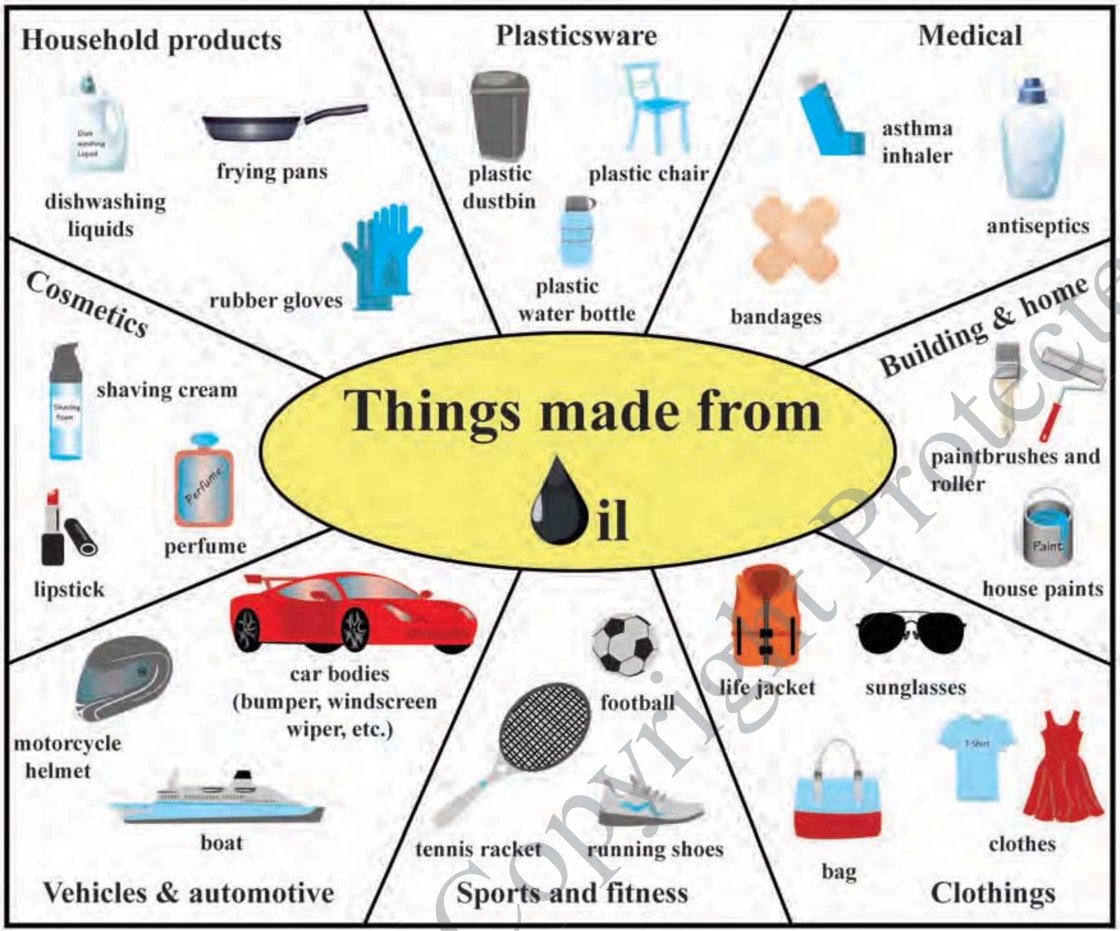
Biogas (Methane or Marsh gas)

Methane gas is formed naturally under a number of different circumstances. Anaerobic bacteria help decomposition of organic matter under geological conditions to produce natural gas. Methane accumulates in coal-mines, where it can cause explosions. Marsh gas, which bubbles up through the stagnant water of marshes, swamps and paddy fields, is also methane. Methane produced in this way contributes to the **'greenhouse effect'**.

Methane is produced from organic waste (biomass) when it decays in the absence of air. This can be exploited as a source of energy. In India and China, biomass digesters are important sources of fuel for villages. Industrialized countries produce large amounts of waste, which is deposited in landfill sites. Biogas forms as the rubbish decays.

Chemistry in Society

- The order for the main energy sources currently used in the world in terms of producing energy is:
  crude oil > coal > natural gas > hydroelectric > nuclear fission > wind > biofuels > solar > geothermal.
- Plant oil, hydroelectric, wind, biofuels, solar and geothermal are all renewable energy sources.
- Natural gas offshore projects in Myanmar are Yadana project, Yetagon project, Shwe Platform project and Zawtika project.



Petroleum products

Natural gas offshore projects in Myanmar

Review Questions

(1) Name each alternative fuel that: (a) is used for cooking and heating systems (b) has the composition of long chain ester (c) is the source of waste organic matter (d) has molecules with 15-20 carbon atoms.

(2) What are the differences between diesel and gasohol?

Key Terms

- **CNG** refers to compressed natural gas (90 % methane) and **LPG** refers to liquefied petroleum gas which is composed of propane and butane.
- **Biodiesel** refers to a vegetable oil- or animal fat-based diesel fuel consisting of long-chain alkyl (methyl, ethyl, or propyl) esters.
- **Transesterification** is the process of exchanging the alkoxy group of an ester compound by another alcohol. These reactions are often catalysed by the addition of an acid or a base catalyst.
- **Biogas** is the mixture of gases produced by the breakdown of organic matter in the absence of oxygen.

EXERCISES

1. Write TRUE or FALSE for each of the following statements. If FALSE, correct it.
   (a) Nowadays, all fossil fuels are not used up rapidly.
   (b) There is a gradual change in the physical properties of the petroleum fractions.
   (c) Hydrogen is a good fuel because it is non-polluting when it burns.
   (d) At present, there is no alternative fuel to fossil fuels.
   (e) Catalytic cracking results in more branched-chain alkanes.

2. Match each of the items given in List A with the appropriate correct item shown in List B.

	List A		**List B**
(a)	coke	(i)	produced from plant oil
(b)	methane	(ii)	formed from waste organic matter
(c)	biodiesel	(iii)	blended fuel from petrol and ethanol
(d)	biogas	(iv)	a reducing agent
(e)	gasohol	(v)	main constituent of natural gas

3. Complete the following sentences by using the words given.
   (a) boiling points, bitumen, fractional distillation, hydrocarbons, crude oil, number of carbon atoms, natural gas, coal, fuel oil
   Fossil fuels are classified into ____, _____ and _____. Petroleum is a mixture of _____. It is separated into different fractions by _____. As the _____ of the fractions increase, the _____ of the fractions increase. The last fraction is _____ and the residue is _____.
   (b) alcohol, ester, transport, petrol-diesel, transesterification, sodium hydroxide, biodiesel
   Plant oils can be converted to _____ by using _____. The plant oils are treated with _____ catalyst, in the presence of 100 % pure ____ at 60 °C. The resultant product is ____. It is a potential substitute for _____ and so it is an alternative _____ fuel.

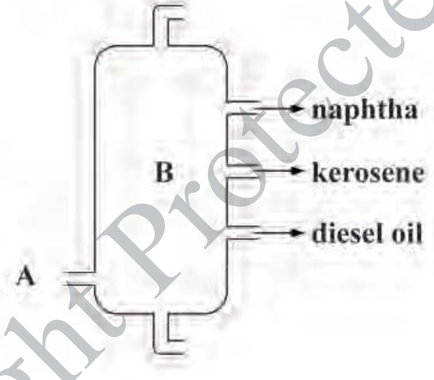
4. Petroleum is considered as a non-renewable fuel while oil from palm trees is considered as a renewable fuel. Explain this statement.

5. (a) Hydrogen is used as a rocket fuel. Why?
   (b) How can you produce hydrogen fuel from (i) natural gas and (ii) water? Write down the chemical equations in words and symbols.

6. Explain the term 'transesterification' in your own words. Describe the application of this reaction.

7. What is meant by the term 'octane number'? Why is it significant?

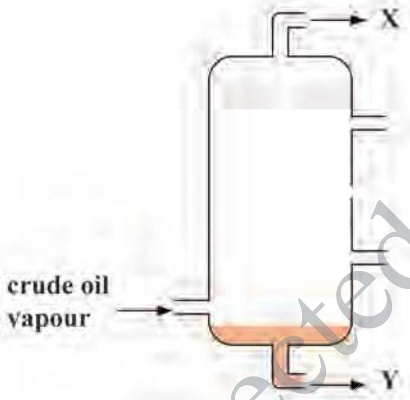
8. The diagram shows how petroleum can be refined.



   (a) What does 'refining petroleum' mean?
   (b) Name the process used to refine petroleum.
   (c) What change of state occurs at A?
   (d) Explain how petroleum is separated at B.
   (e) State (i) two similarities; (ii) two differences between naphtha and diesel oil.
   (f) Name two fuels, suitable for cars, which are not obtained from petroleum.
   (g) Among three fractions, which fraction has the lowest boiling point and which one has the highest boiling point?
   (h) Among three fractions, which fraction contains the smallest molecules and which one contains the biggest molecules?

9. Petroleum is a mixture of hydrocarbons. The different hydrocarbons have different boiling points and petroleum has to be separated into its various components before the individual components can be used.
   (a) How would you explain the term 'hydrocarbon'?
   (b) Name the separation method used to separate petroleum in oil refineries. What physical property of liquids makes separation by this method?
   (c) Suggest the name of a petroleum fraction that would be suitable for each of the following purposes:
      (i) seal cracks in the concrete tanks
      (ii) boil a beaker of water in the laboratory
      (iii) protect a wooden furniture
      (iv) oil the sewing machine to reduce friction

10. Cracking is a process that split larger hydrocarbons into smaller ones.
   (a) Give two reasons why an oil company might want to crack a hydrocarbon.
   (b) Give the conditions under which cracking is carried out.
   (c) A molecule of the hydrocarbon $C_{11}H_{24}$ was cracked to give two molecules of ethene (C_2H_4) and one other molecule. Write a balanced chemical equation for the reaction which took place.
   (d) Write a chemical equation for an alternative cracking reaction involving the same hydrocarbon $C_{11}H_{24}$.

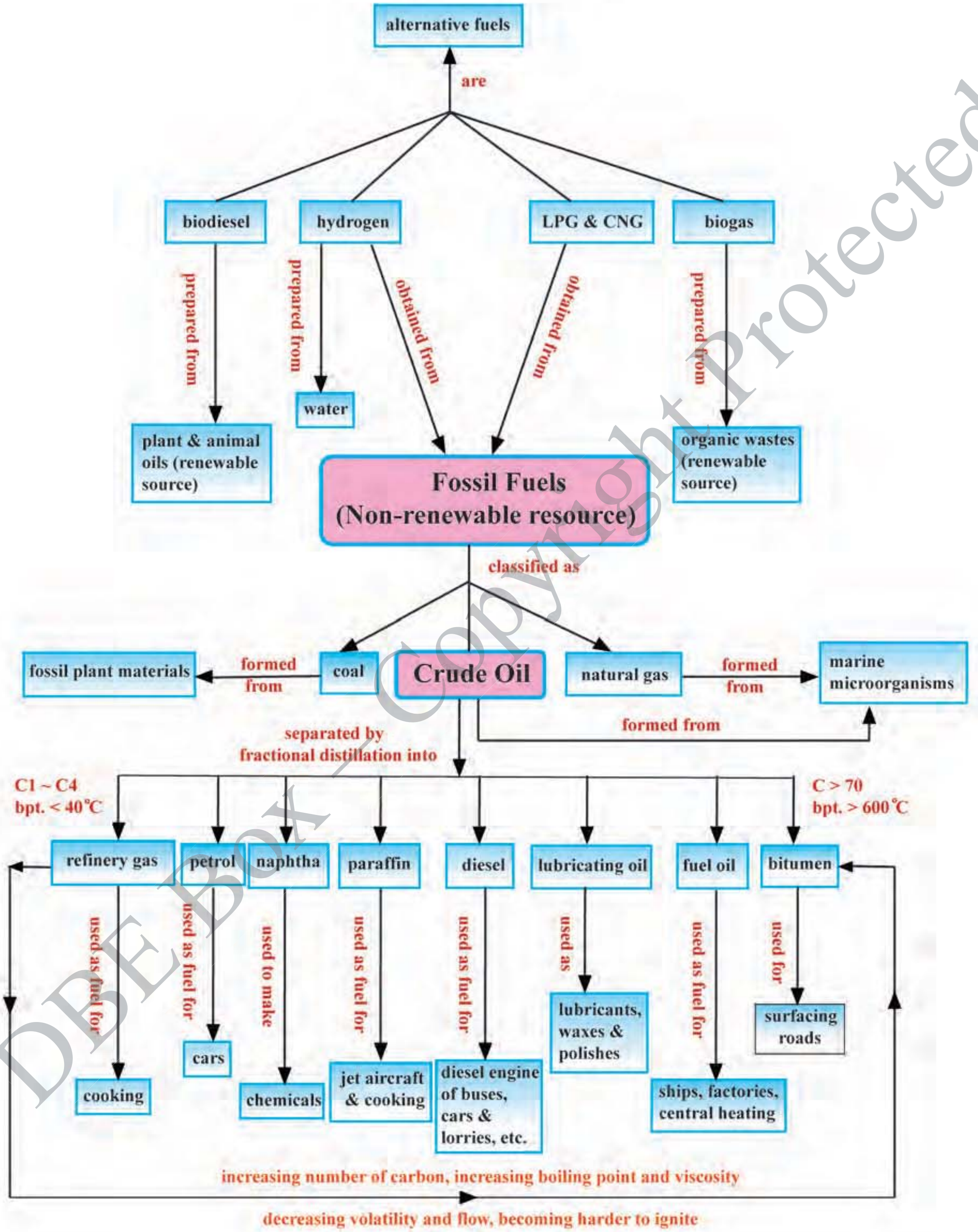


11. The diagram given represents the process of fractional distillation of crude oil.
Which of the following statements about fractions **X** and **Y** is correct?
Give reason for your answer.
    - (a) **X** burns more easily than **Y**.
    - (b) **X** has a higher boiling point than **Y**.
    - (c) **X** is used for making road surfaces.
    - (d) **Y** is the lighter fraction compared with **X**.

12. Consider the following petroleum fractions: diesel oil, kerosene, naphtha, petroleum gas, bitumen.
    - (a) Place the above fractions in order of decreasing volatility from most volatile to least volatile.
    - (b) Of the fractions given above,
        - (i) which contains hydrocarbon with more than 20 carbon atoms per molecule?
        - (ii) which contains propane (C_3H_8)?
        - (iii) which is used as a starting material for making plastics?

13. A hydrocarbon, **Q**, was found to contain 82.8 % carbon by mass.
    - (a) What is the empirical formula of **Q**?
    - (b) The molecular formula of **Q** is C_xH_{10}. In which petroleum fraction would **Q** be found? Explain your answer.

CHAPTER REVIEW
(Concept Map)



GLOSSARY

Acid	a compound that dissolves in water to produce hydrogen ions, H^+
Acidic oxide	a non-metallic oxide which reacts with basic oxide to produce salt
Acid rain	the rain formed when acidic air pollutants mainly sulphur dioxide and nitrogen dioxide react with water in the air (atmosphere)
Air pollution	the condition in which air contains a high concentration of air pollutants that may harm living things and also damage non-living things
Alkali	a base which is soluble in water producing OH^- ions
Allotropes	Two or more forms of an element that occur in the same physical state but different in properties
Alloy	a substance made by combining two or more metallic elements, especially to give greater strength or resistance to corrosion
Amphoteric oxide	a metallic oxide which possesses both basic and acidic properties
Anion	a negatively charged ion
Atom	the smallest particles into which an element can be divided
Atomic number (Z)	the number of protons in the nucleus of an atom of an element
Avogadro's constant	the number (6.02×10^{23}) of entities or a stated type of particles (atoms, ions or molecules) in a mole of those substances
Base	a chemical compound that combines with an acid to form a salt and water
Basic oxide	a metallic oxide which reacts with acid to form salt and water
Binary compounds	compounds formed by the combination of two elements
Biodiesel	a vegetable oil- or animal fat-based diesel fuel consisting of long-chain alkyl (methyl, ethyl, or propyl) esters
Biogas	the mixture of gases produced by the breakdown of organic matter in the absence of oxygen
Boiling point	the temperature at which the vapour pressure of the liquid is equal to the atmospheric pressure of the surrounding
Brownian motion	the continuous random movement of small particles suspended in a gas or liquid, which arises from collisions with the gas or liquid particles, e.g., the motion of pollen grains on still water, movement of invisible dust in a room
Catalyst	a substance that speeds up the chemical reaction without getting consumed
Catalytic cracking	a process used to split long chain alkanes into shorter alkanes and alkenes in the presence of catalyst under pressure and high temperature
Cation	a positively charged ion

Centrifugation	a technique used for the separation of particles from a solution according to their size, shape, density, viscosity of the medium and rotor speed
Chemical change	a change in which one or more new substances are formed
Chromatography	a separation method of mixed substances that depends on the speed at which they move through special media, or chemical substances
CNG	Compressed Natural Gas (90 % methane)
Complete electronic structure	the arrangement of all the electrons of an atom of the element in appropriate sub-shells
Compound	a substance containing two or more different elements chemically joined together in a fixed ratio
Compound oxide	the combination of two different oxides of the same element
Concentrated acid	an acid solution which contains the pure acid or predominantly large proportion of the acid
Condensation	the change from a gaseous state to its liquid state
Covalent bond	a bond formed by sharing of electrons between two atoms by weak intermolecular force of attraction
Crude oil	a mixture of many different hydrocarbon molecules
Crystallisation	a process by which a chemical is converted from a liquid solution into a solid crystalline state
Decantation	a process to separate mixtures of solid and liquid or two immiscible liquids to settle and separate by gravity
Degree of hardness of water	the number of parts of mass of $CaCO_3$ (calcium carbonate), equivalent to various calcium and magnesium salts present in one million parts by mass of water (ppm)
Deposition	the direct solidification of a vapour by cooling; the reverse of sublimation
Diatomic molecules	molecules composed of only two atoms of same or different elements
Dilute acid	an acid solution which contains a relatively small amount of the acid
Dissociation reaction	a chemical reaction in which a compound breaks apart into two or more parts
Electron affinity	the energy released when an electron is added to a gaseous atom to form a gaseous ion
Electronegative elements	non-metals which tend to gain electrons and form negative ions
Electronic structure	the distribution of electrons in an atom of an element
Electropositive elements	metals which tend to lose electrons and form positive ions
Element	a substance that cannot be broken down into other simpler substances through chemical means

Empirical formula	formula which shows the simplest whole number ratio of atoms in a compound
Engine knocking	the sharp sounds caused by combustion of some of the compressed air-fuel mixture in the cylinder.
Essential electronic structure	the representation of the arrangement of valence electrons of an atom of the element in appropriate sub-shells
Evaporation	the process of a substance in a liquid state changing to a gaseous state due to an increase in temperature and / or pressure
Filtration	a method for separating an insoluble solid from a liquid
Fossil fuels	the fuels consisting of coal, petroleum (crude oil) and natural gas
Fractional distillation	a method for separation of a liquid mixture into fractions with different boiling points (and hence chemical composition) by means of distillation, typically using a fractionating column
Freezing	the process in which a liquid becomes sufficiently cold to change into a solid
Freezing point	the temperature at which a liquid becomes a solid
Fuel	a substance that releases sufficient energy to do work as it undergoes a chemical change
Global warming	the increase in the Earth's average temperature due to the built-up of greenhouse gases in the atmosphere
Hard water	water which will not readily form a permanent lather with soap
Heterogeneous mixture	one that is non-uniform, and where the different components of the mixture can be seen (The components separate, and the composition varies.)
Homogeneous mixture	one in which the composition of its components are uniformly mixed throughout (The components cannot be seen separately on visual or microscopic examination.)
Hydrocarbon	any of a class of organic chemical compounds composed only of the elements carbon (C) and hydrogen (H)
Indicator	a substance that has different colours in acidic and alkaline solutions
Ion	a charged particle
Ionic bond	a bond formed by the complete transfer of an electron or electrons from one atom to another resulting in the formation of cations and anions. These oppositely charged ions are held together by a force of electrostatic attraction
Ionisation energy	the amount of energy required to remove an electron from a gaseous atom to form a gaseous ion
Isobars	the atoms with same mass number but different atomic numbers
Isotopes	atoms of the same element that have the same number of protons but different number of neutrons (or) the atoms of the same element with different masses

Law of conservation of mass	the law that states that the total mass of the reactant(s) equals to the total mass of the product(s)
LPG	Liquefied Petroleum Gas which is composed of propane and butane
Magnetic separation	a method used to separate the components of a mixture when at least one of them is magnetic in nature
Matter	a substance made up of tiny particles, and has mass and takes up space. Three common states of matter are solid, liquid and gas.
Mass number (A)	the sum of the number of protons and neutrons or the total number of nucleons in the nucleus of an atom of an element
Melting point	the temperature at which a solid changes to a liquid state at one atmospheric pressure
Metalloids	the elements that have the properties of both metals and non-metals
Mixture	a combination of more than one substance, where these substances are not bonded to each other (It consists of two or more substances which may be present in any proportion by weight. The constituents of the mixture do not combine chemically.)
Molar mass	the mass of one mole of a substance
Molar volume of a gas	a volume of 24 dm^3 or 24,000 cm^3 at room temperature and pressure (r.t.p.) for one mole of gas
Molecular formula	formula that shows the total number of atoms of each element present in one molecule or one formula unit of the compound
Molecule	the simplest unit of a chemical substance, usually a group of two or more atoms
Neutral oxide	an oxide that does not react with either acids or bases
Neutralisation	the reaction between an acid and a base to form a salt and water only
Non-renewable fuels	fuels which take millions of years to form and which are used up at a rapid rate
Nucleon number	the total number of protons and neutrons in the nucleus of its atom
Octane rating	(octane number) a measure of a fuel's ability to resist 'knock'
Octet rule	a rule which states that in forming a chemical bond, atoms gain, lose or share electrons in such a way to attain the stable electronic structures of the noble gases, i.e., to have eight electrons in the outermost shell
One mole of a substance	the amount of substance that has the same number of particles (atoms, molecules, etc.) as there are atoms in exactly 12 g of ^{12}C
Organic matter	a substance that is made up of undecomposed and partially decomposed residue of plant and animal tissues of living and dead microorganisms

Oxidation number	the combining capacity of the element and also indicates the positive and negative nature of its atoms in the compounds
Periodic Table	a list of chemical elements arranged in order of atomic number in rows, so that elements with similar electronic structures (and hence, similar chemical properties) appear in vertical columns
Permanent hardness of water	hardness of water caused by dissolved calcium or magnesium chloride and sulphate which cannot be removed by boiling
Permutit method	a process in which hard water containing calcium or magnesium salt is passed through a layer of sodium zeolite, the calcium is abstracted and the corresponding sodium salt passes in solution.
Peroxide	an oxide that reacts with an acid to produce salt and hydrogen peroxide
pH	a measure of the acidity or alkalinity of a solution
Physical change	a change in which no new substances are formed (There may be a temporary change in colour, temperature and state of the substances but no new substances are formed.)
Polyatomic molecules	molecules composed of three or more atoms of same or different elements
Product	the substance that is produced in the reaction
Reactant	the substance that takes part in the reaction
Relative atomic mass	the average mass of one atom of that element compared to one twelfth the mass of one atom of carbon-12
Relative formula mass	the sum of the relative atomic masses of all the atoms in the formula
Relative molecular mass	the mass of one molecule of a substance compared to one twelfth the mass of one atom of carbon-12
Renewable fuels	fuels produced from renewable resources
Salt	a substance produced from the reaction between an acid and a base or a metal
Saturated solution	a solution in which no more solute will dissolve at the given temperature, in the presence of excess solute
Simple distillation	a procedure by which two liquids with different boiling points can be separated
Soft water	water which readily gives a permanent lather with soap
Soil texture	a measure of the relative propotion of sand, silt and clay separates (particles) found in the soil
Solubility	the mass in grams of the substance which will saturate 100 g of water, at given temperature
Solute	a substance which dissolves in a solvent to give a solution

Solution	a clear homogeneous mixture obtained when a substance dissolves in a solvent (In a solution the solute is uniformly distributed throughout the solution.)
Solvent	a substance, mostly liquid, in which another substance dissolves to give a homogeneous mixture
STP	standard temperature, 0 ^{o}C or 273 K and standard pressure, 760 mmHg or 1 atmosphere
Strong acid	an acid that completely dissociates in water and gives H^+ ions
Strong base	a base that completely dissociates in water producing OH^- ions
Sublimation	the change of solid state directly into gaseous state without melting
Supersaturated solution	a solution that retains more solute than that required to saturate the solution at room temperature
Temporary hardness of water	hardness of water caused by dissolved calcium hydrogen carbonate which is removed by boiling
Ternary compounds	the compounds formed by the combination of three elements
Thermal decomposition	the breaking down of a chemical compound, caused by heat
Transesterification	the process of exchanging the alkoxy group of an ester compound by another alcohol
Triatomic molecules	molecules composed of only three atoms of same or different elements
Unsaturated solution	a solution in which more of the solute can dissolve at the given temperature
Valence	the number of electrons in the outermost shell when the number of electrons in the outermost shell is 4 or less, or equals to 8 minus number of electrons in the outermost shell when the number of electrons in the outermost shell is greater than 4.
Vaporisation	the process that occurs when a chemical or element is converted from a liquid to a vapour
Water hardness	the amount of dissolved calcium ions, magnesium ions or both in the water
Weak acid	an acid that partially dissociates in aqueous solution and gives H^+ ions
Weak base	a base that partially dissociates in water producing OH^- ions
Zeolites	microporous, aluminosilicate minerals commonly used as commercial adsorbents and catalysts

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