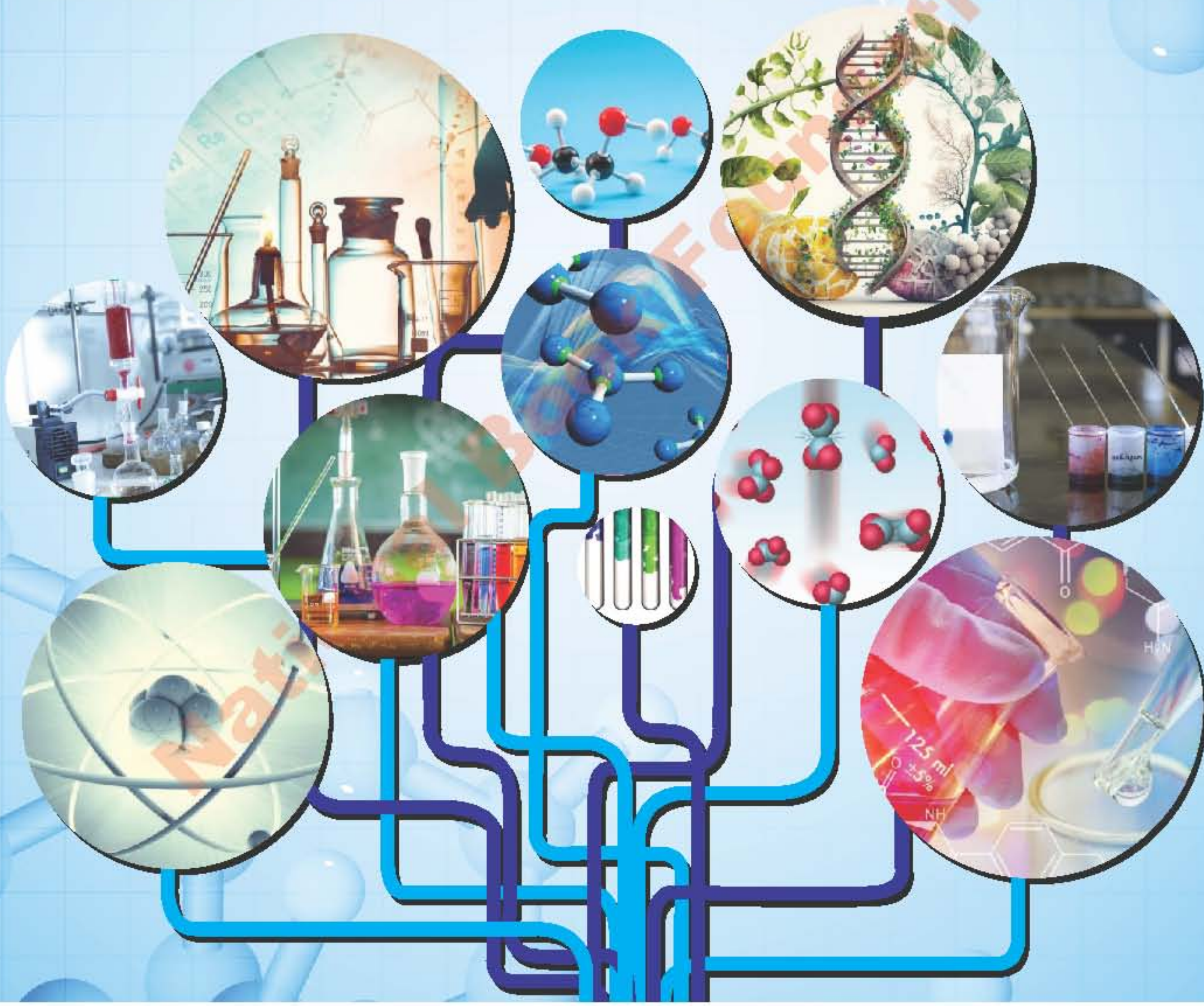
Model Textbook of

CHEMISTRY

Grade 9

Based on National Curriculum of Pakistan 2022-23



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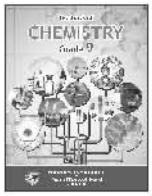

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PREFACE

In a historic footstep, the national curriculum of Pakistan 2022-2023 has introduced a new era for schooling in the country, This is the first-ever core curriculum in the 75-year history of Pakistan. It is in line with the protected right to school education by Article 25-A.

Chemistry might be a difficult subject for someone, but it holds significance for those who embrace a systematic approach to understanding its concepts.

This new Textbook has been developed as a model Textbook for Pakistan. The book consolidates critical thinking methodologies, guiding scientific reasoning, and thinking abilities. The book incorporates problem-solving strategies, which will guide students toward analytical thinking and skills. These skills would be invaluable for both academic as well as practical life.

The book also inspires concept assessment exercises in every unit, which have been designed to evaluate acquired knowledge and promote critical thinking and analyzing data..

One of the book`s distinctive features is the key points at the end of each unit, which serve as a quick reference to reinforce the salient features of each unit.

Dr. Raja Mazhar Hameed

Managing Director

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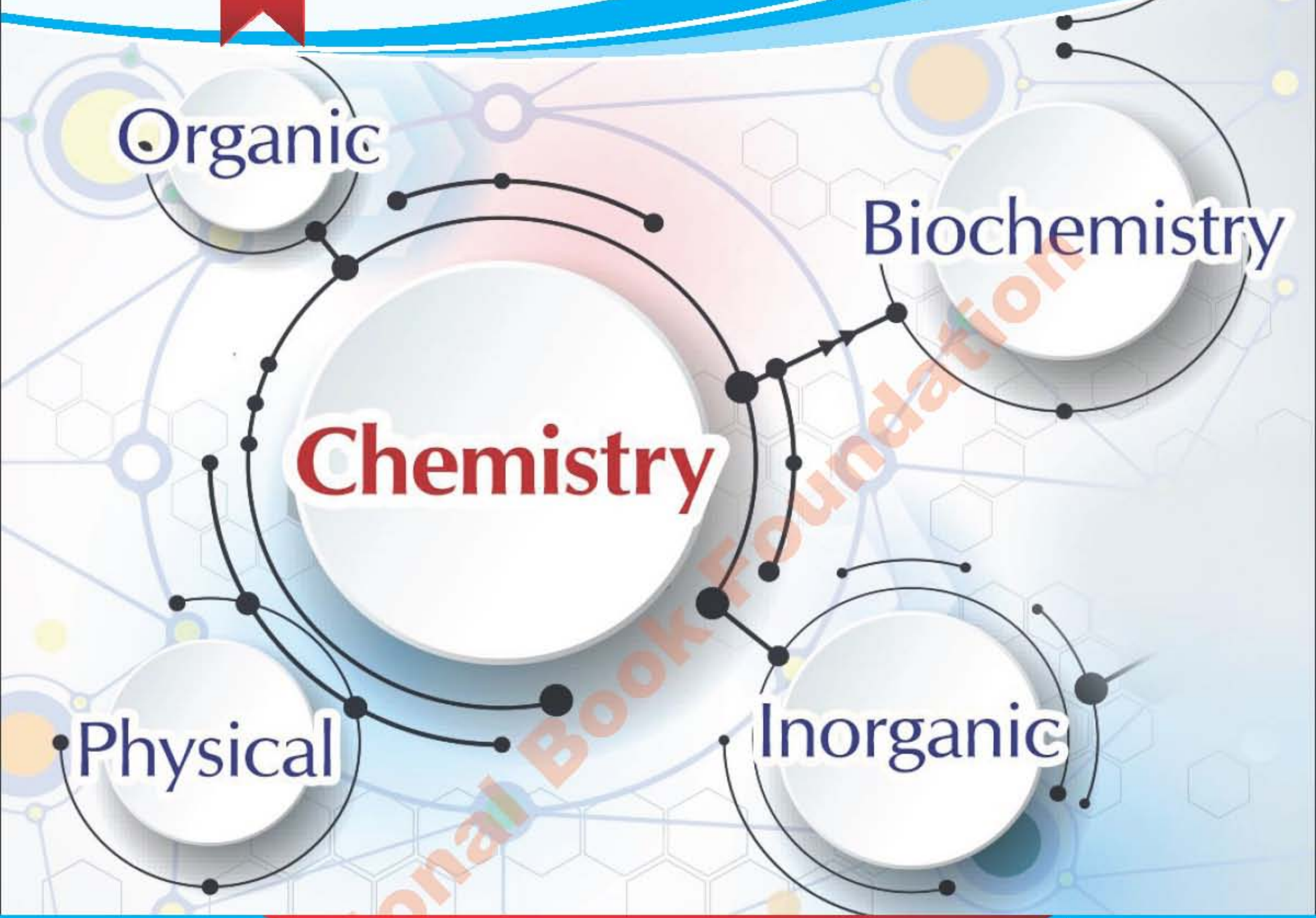
UNIT 01



NATURE OF SCIENCE IN CHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define chemistry as the study of matter, its properties, composition, and its interactions with other matter and energy.
- Explain with examples that chemistry has many sub-fields and interdisciplinary fields.
- Formulate examples of essential questions that are important for the branch of chemistry.
- Differentiate between 'science' 'technology' and 'engineering' by referring to examples from the physical sciences.

1.1 DEFINITION OF CHEMISTRY AND ITS INTERACTION WITH OTHER MATTER AND ENERGY:

Chemistry is defined as the science that investigates the materials of the universe and the changes that these materials undergo. Chemistry deals with the composition, structure, properties, behavior, and changes of matter and energy. Understanding the fundamental concepts of chemistry help to explain natural phenomena and apply them to the formation of new substances, drugs, and technologies.

DO YOU KNOW?

- How green chemistry is helpful in understanding and reducing pollution?
- Green Chemistry is the model of chemical products and processes that reduce the use of hazardous substances.

1.2: BRANCHES OF CHEMISTRY

Chemistry is a diverse field of study, surrounding numerous sub-fields and interdisciplinary areas.

1. Organic Chemistry

Organic chemistry is a branch of chemistry that deals with substances containing carbon (except carbonates, bicarbonates, oxides, and carbides).

2. Inorganic Chemistry

Inorganic chemistry is a branch of chemistry that deals with elements and their compounds except organic compounds.

3. Physical Chemistry

Physical chemistry is the branch of chemistry that deals with laws and theories to understand the structure and changes of matter.

4. Analytical Chemistry

Analytical chemistry is a branch of chemistry that deals with the methods and instruments for determining the composition and properties of matter.

5. Biochemistry

The branch of chemistry that deals with physical and chemical changes that occur in living organisms is called biochemistry.

6. Environmental Chemistry

Environmental chemistry is the branch of chemistry that deals with the study of chemical and toxic substances that pollute the environment and their adverse effects on human beings.

7. Industrial Chemistry

Industrial chemistry is the branch of chemistry that deals with the large-scale production of chemical substances.

8. Medicinal Chemistry

The branch of chemistry deals with the study of the interaction between drugs and biological targets, as well as the development of new medicinal agents.

9. Polymer Chemistry

The branch of chemistry that focuses on the study of polymers, their types, properties, uses, importance, and types of polymerizations is called polymer chemistry. Examples of synthetic polymers include nylon bearings, plastic bags, polyethylene cups, polyester, Teflon coated cook ware, and epoxy glue etc.

10. Geochemistry

Geochemistry is the branch of chemistry that deals with the study of chemical composition, distribution, and transformation of elements and compounds in the Earth's crust, such as rocks, minerals, soils, water, and the atmosphere.

11. Nuclear Chemistry

The branch of chemistry that deals with the changes that occur in atomic nuclei is called nuclear chemistry.

12. Astrochemistry

Astrochemistry is a branch of chemistry that deals with the study of chemical processes and reactions that occur in astronomical environments, such as stars, planets, comets, and interstellar space.

1.3 EXAMPLES OF ESSENTIAL QUESTIONS THAT ARE IMPORTANT FOR THE BRANCHES OF CHEMISTRY

Some essential questions for various branches of chemistry that can help enhance understanding are as follows:

Physical Chemistry

1. What is the structure of an atom, and how does it influence chemical behavior?
2. How do different types of chemical bonds (ionic, covalent, metallic) form and function?

Organic Chemistry

1. Why carbon is considered the backbone of organic compounds?
2. What are the major functional groups in organic molecules, and how do they affect chemical properties?

Inorganic Chemistry

1. What distinguishes inorganic compounds from organic compounds?
2. How does Periodic table helps to organise elements?

Analytical Chemistry

1. How are analytical methods used to identify and quantify chemical substances?

Biochemistry

1. How do biomolecules such as carbohydrates, proteins, nucleic acids, and lipids contribute to the structure and function of living organisms?

Environmental Chemistry

1. How do human activities contribute to air pollution, and what are the consequences for the environment?
2. What role do greenhouse gases play in climate change, and how can we mitigate their effects?

Medicinal Chemistry

1. How are drugs designed and developed for specific therapeutic purposes?

Polymer Chemistry

1. What are polymers, and how do their structures affect their properties?

Geochemistry

1. How do geological processes influence the distribution of elements in the Earth's crust?

Nuclear Chemistry

1. How do nuclear reactions differ from chemical reactions, and what are their applications?
2. What is the role of radioisotopes in medicine and industry?

Astronomy

1. What types of reactions occur in astronomical environments?

These questions can serve as a foundation for exploring the key concepts within each branch of chemistry.

1.4: DAILY LIFE APPLICATIONS OF CHEMISTRY

Organic Chemistry

To treat diseases, organic chemists synthesize new medicines that interact with specific targets like proteins or enzymes.

Inorganic Chemistry

Lithium-ion (Li-ion) batteries are used as rechargeable batteries for electronics, toys, wireless headphones, handheld power tools, small and large appliances, electrical storage devices, and electric vehicles.

Analytical Chemistry

Forensic chemistry is the application of analytical chemistry. It involves the examination of physical traces, such as body fluids, bones, fibers and drugs. It can be used to identify an unknown compound. For example drugs are often found in various colored powders and are analyzed to determine their content.

Physical Chemistry

Physical chemistry is a part of our everyday life. The batteries in our vehicles are built on the principle of electrochemistry.

Environmental Chemistry

Environmental chemistry is used to protect water that has been poisoned by soil, and dust by using different methods e.g., sedimentation, filtration, and disinfection.

1.5: 'SCIENCE' 'TECHNOLOGY' AND 'ENGINEERING'

Science

Science is the systematic process of constructing and organizing knowledge about the universe. Thus, science seeks to understand the natural world. For example, chemists seek to understand the behaviour and properties of materials, chemical reactions, and the fundamental principles that control the behaviour of matter.

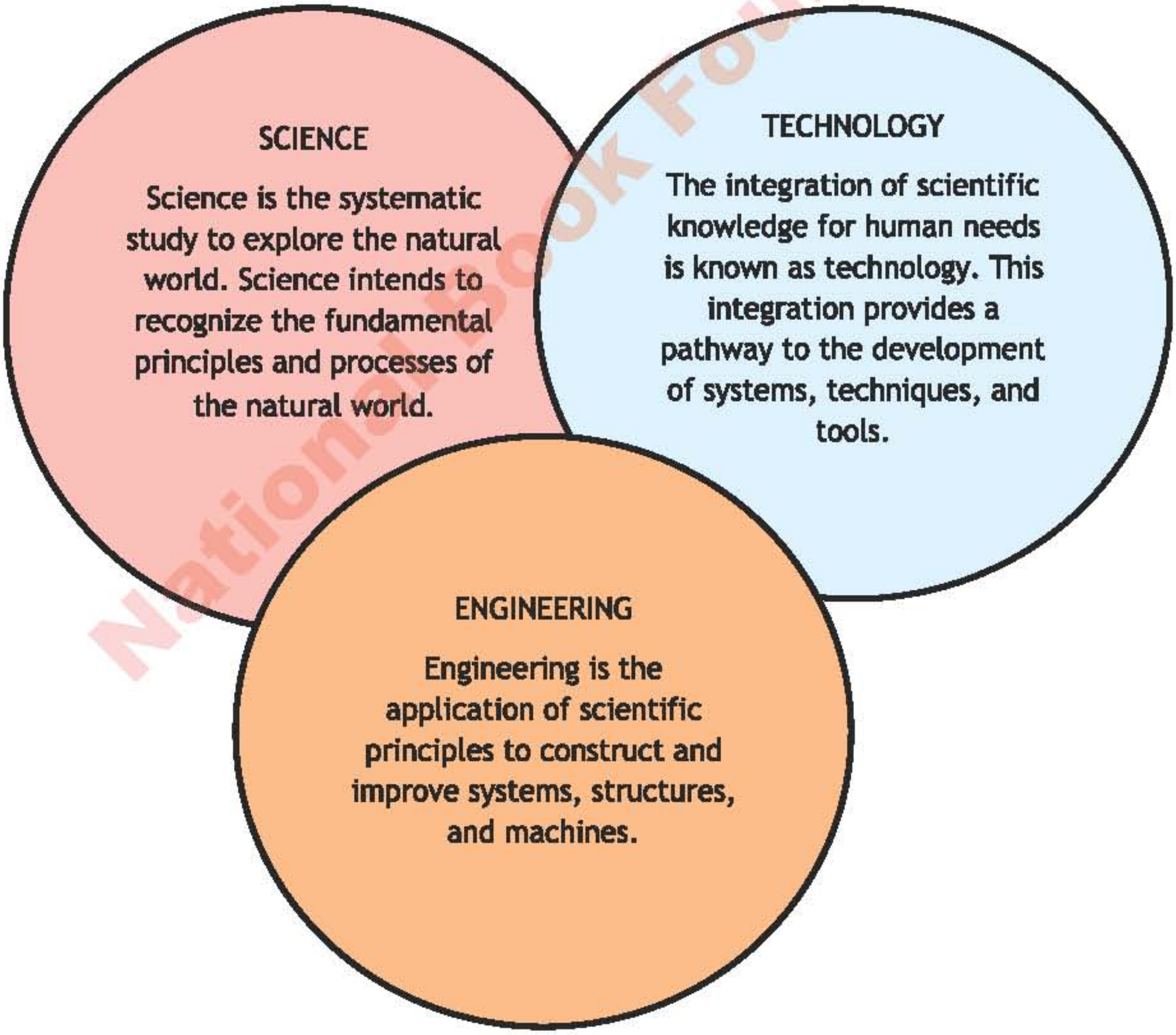
Technology

Technology is the process of applying scientific knowledge to practical applications, resulting in the creation of tools, machines, and systems that enhance our lives.

Science and technology play a major role in the field of chemistry by providing tools, machines, techniques and methods which can help in discovery and ,development of new materials .These also help in improving quality of products .Technology has revolutionized the field of chemistry, making research and applications more efficient. It has enabled chemists to more effectively analyze and identify substances. Their work is beneficial for chemists working in pharmaceutical and other chemical industries.

Engineering

Engineering is the use of science and mathematics to design and construct systems, structures, and tools for various processes. Chemical engineers develop and design manufacturing processes for the production of chemicals , fuels, food, medicines, polymers , detergents, paper etc. They often work to maximize productivity and product quality.



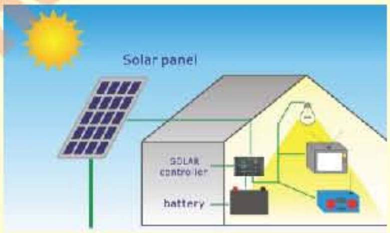
1.6: APPLICATIONS OF SCIENCE AND TECHNOLOGY AND ENGINEERING

Let's take a look at how science, technology and engineering work together to solve problems in real-world situations. For example:

Example 1.1: Investigating rusting of Iron.

Imagine trying to figure out why a bike or car will rust over time. Scientists could investigate the chemical reactions that occur between iron, water and oxygen that cause rust to form. Experiments could be conducted to understand the factors that influence this process and help develop strategies to prevent rust.

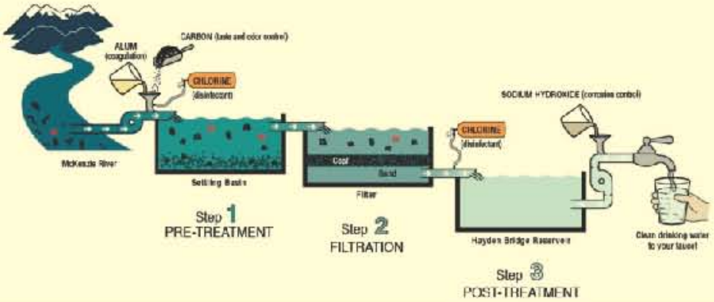
Example 1.2: Harnessing Solar Energy



Scientists may study the principles of photovoltaic cells to understand how sunlight can be converted into electricity. Technologists can develop solar panels based on the scientific principles discovered. Engineers play their role in designing and implementing large-scale renewable energy systems. For instance, an electrical engineer might design the wiring and connections of a solar power plant, a civil engineer could be involved in designing the infrastructure. In this example, science helps us understand the underlying principles of converting sunlight into electricity. Technology transforms this knowledge into practical applications, such as solar panels and energy storage systems. Engineering takes these technologies and implements them on a larger scale. Together, science, technology, and engineering contribute to the development and utilization of sustainable energy sources.

Example 1.3: Designing a Water Filtration System

let's look at how engineers design water filtration systems. Chemical engineers can help develop processes to remove contaminants from water while mechanical engineers design physical components. Together, they create a solution to clean drinking water for a community. From understanding chemical reactions to using technological devices to solving practical problems through engeneering.



Example 1.4: Organic Chemistry in Action

How do you make french fries. The oil used to fry potatoes contains carbohydrates, which are organic molecules. Scientists study carbohydrates to learn more about how they work, so food technologists extract oil from seeds. Chemical engineers design oil production equipment and processes so that oil is produced efficiently and safely for cooking.

Example 1.5: Plastic Bags

Think about the science behind plastic bags. Scientists study the small building blocks known as monomers. When monomers combine, they form long chains known as polymers. One of those long chains is polyethylene, which is one of the many polymers found in plastic bags! Engineers and technicians use these discoveries to create bags that are durable, flexible, and easy to make.

These examples demonstrate how science, technology, and engineering work together in various aspects of our daily lives. Whether it's understanding chemical reactions, using technological devices, or solving practical problems through engineering solutions, these concepts are interconnected and contribute to advancements that impact the world around us.

KEY POINTS

- Chemistry is the study of matter around us.
- The branch of chemistry deals with carbon compounds (except bicarbonates, carbonate oxides, and carbides.
- The branch of chemistry that deals with the elements and their compounds except organic compounds is called inorganic chemistry.
- Industrial chemistry is concerned with the large-scale production of chemical substances.
- The branch of chemistry that deals with the laws and theories to understand the structure and changes of matter is called physical chemistry.
- Science is defined as the study of nature.
- Technology is the application of science.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which branch of chemistry is the study of elements and their compounds except for organic compounds?

(a) Physical Chemistry (b) Organic Chemistry
(c) Inorganic Chemistry (d) Geochemistry Chemistry

(ii) Which branch of chemistry helps to protect water that has been poisoned by soil?

(a) Environmental Chemistry (b) Organic Chemistry
(c) Inorganic Chemistry (d) Geochemistry Chemistry

(iii) Which area of Chemistry improves to gauge the behavior of pollutants and develop techniques for pollution control?

(a) Analytical Chemistry (b) Organic Chemistry
(c) Environmental (d) Geochemistry

(iv) The branch of chemistry that helps to treat diseases, organic and to synthesize new medicines.

(a) Physical (b) Organic
(c) Inorganic (d) Environmental

(v) The branch of science helps to understand chemical products and processes that reduce the use of hazardous substances:

(a) Analytical Chemistry (b) Physical chemistry
(c) Green Chemistry (d) astrochemistry

(vi) To identify the concentration of a particular solution through titration is and application of

(a) Astrochemistry (b) Analytical Chemistry
(c) Geochemistry (d) Organic chemistry

(vii) The batteries in our vehicles are built on the principle of electrochemistry. It is the application of:

(a) Astrochemistry (b) Analytical Chemistry
(c) Organic chemistry (d) Physical chemistry

(viii) The branch of chemistry that is concerned with the large-scale production of chemical substances is:

(a) Industrial chemistry (b) Physical chemistry
(c) Inorganic chemistry (d) Environmental Chemistry

(ix) The branch of chemistry that focuses on the study of polymers, their types, properties, uses is called:

(a) Industrial Chemistry (b) Polymer chemistry
(c) Organic Chemistry (d) astrochemistry

(x) The study of the interaction between drugs and biological targets, as well as the development of new medicinal agents.

(a) Organic chemistry (b) Medicinal chemistry

(c) Inorganic chemistry (d) Environmental Chemistry

2. Give short answer.

(i) How does chemistry help a doctor to know about the chemical nature of medicine?

(ii) In what ways does technological innovation help to understand the development of new materials?

(iii) Differentiate between geochemistry and astrochemistry.

(iv) With the help of an example correlated the use of science, technology, and engineering.

(v) With the help of the Venn diagram compare and contrast organic and inorganic chemistry.

(vi) What are the uses of nuclear chemistry?

3. Define chemistry and its interactions with other matter and energy.

4. Describe the applications of inorganic chemistry and its importance in our daily lives?

5. With the help of few examples highlight the relation between science, technology and engineering.

6. Evaluate the role of chemistry in environmental science.

7. How does geochemistry help us to solve the problems such as pollution and climate change?

8. How is organic chemistry applied in medicines, biochemistry and industrial science?

PROJECT

1. Draw figure of a tree showing different branches of chemistry.

2. Composting is a great way to recycle materials that might be thrown into landfill. It takes years to decompose them. Make an indoor composter and determine how readily different materials decompose.

UNIT 02

MATTER

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define matter as a substance having mass and occupying space.
- State the distinguishing macroscopic properties of commonly observed states of solids, liquids, and gases in particular density, compressibility and fluidity.
- Identify that state is a distinct form of matter (examples could include familiarity with plasma, intermediate states and exotic states e.g. BEC or liquid crystals).
- Explain the allotropic forms of solids (some examples may include diamond, graphite, and fullerenes).
- Explain the differences between elements, compounds, and mixtures.
- Identify solutions, colloids and suspensions as mixtures and give an example of each.
- Explain the effect of temperature on solubility and formation of unsaturated and saturated solutions

INTRODUCTION

The study of chemistry revolves around the study of matter which is all around us; not only is the entire world made up of matter but so are we, so are the objects that we use. From this we can derive the definition of matter:

Anything that has mass and occupies space is called matter. This makes air, water, rocks, and even people are examples of matter. Different types of matter can be described by their mass. Matter is itself composed of the atom. The atom is the building block of all matter and it is the various combinations of these atoms that make up all the matter that we see around us. You may ask yourself how the book you are reading and the water you are drinking are both matter. They neither look nor feel nothing alike. So how can they both fall into the definition of matter? From there we reach the conclusion that there are states of matter which differ from each other in the way that the atoms that make them up are arranged:

2.1 STATE OF MATTER

There are four states of matter

1. **Gas**
2. **Liquid**
3. **Solid**
4. **Plasma**

Each state is a distinct form of matter.

1. States of matter are the different forms in which matter can exist. These are solids, liquids, gases, and plasmas. These states are determined by the arrangement and movement of particles and the strength of intermolecular and atomic forces.
2. Energy can change matter into different states. For example, solids become liquids or gases when heated. At very high temperatures or when subjected to a strong electric field, the gas transforms into plasma. Under normal conduction, most substances remain in one distinct state: solid, liquid, or gas. Temperatures and energy levels on the Earth are not sufficient to ionize atoms and create plasma.
3. When heated, some crystalline solids turn into cloudy liquids that completely dissolve. This cloudy state is called liquid crystal. Liquid crystal states have many properties of liquids and some properties of solids. This form exists within a certain temperature range. When heated further, the state of the liquid crystal changes to a transparent liquid.
4. Furthermore, there are other states such as Bose-Einstien Condensates(BEC) which is defined as the state of matter in which separate atoms cooled to temperatures very close to absolute zero. BEC is observable under extreme conditions of cold temperature. Superfliud and superconductors are the two main materials which contain BEC.

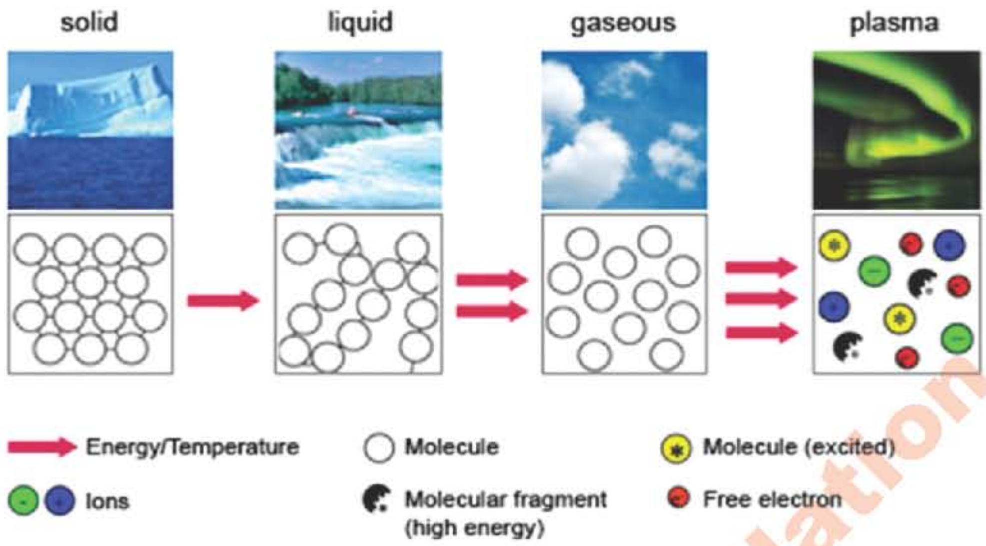


Figure 2.1: State of MAtter

Macroscopic properties that can be visualized by the naked eye and we can take measurements easily. Some common examples of macroscopic properties of matter include density, fluidity, compressibility.

Table 2.1: Properties of different states of matter

State of matter	Gas	Liquid	Solid
Density	Low density at normal condition due large spaces between molecules	High density at normal condition	High density at normal condition
Compressibility	Very compressible because of large empty spaces	moderately compressible	not compressible
Fluidity	Can flow	Can flow	Can not flow

Have you ever boiled water on a stove? What do you observe when the water heats up? Bubbles form and the water turns into the gas. This tells us a very important fact about the states of matter. Though the states of matter are distinct and are easily distinguishable from the other, through physical techniques we can convert one state of matter into the other. Physical techniques are techniques where we manipulate the physical aspects of matter such as the temperature or pressure. However, the chemical composition of matter stays the same.

2.2 ELEMENTS, COMPOUNDS AND MIXTURES

Matter can be described with both physical properties and chemical properties. Matter can be classified as

1 **Pure substance**
  a) Element
  b) Compound
2 **Mixture**
  a) Homogeneous
  b) Heterogeneous mixtures are
    i) Colloid
    ii) Suspension

Earlier, we talked about the atom and how atoms make up all of matter. Same types of atoms are called *elements*. An element consists of atoms that have the same atomic number also known as the proton number. This is the simplest form of matter which cannot be broken down through chemical means. While a physical change alters the physical properties of a substance, a chemical change forms a new substance completely.

***Element*:** the simplest form of matter made up the same type of atoms

So we have learnt that matter is made up of atoms and the atoms that have the same proton number are called elements. The combination of these different elements makes up the diversity of objects we see around us. When two or more elements chemically combine, meaning undergo a chemical reaction to form a new substance, this is called a compound. As this is a completely new substance, it is completely different from the elements that were used to make it.

Compound: A substance formed when two or more different atoms chemically combine.

Mixtures are the physical combinations of substances. A mixture that does not contain the same types of particles. If you were to examine the chemical composition of the particles in a mixture, the particles would be chemically different from each other. Tea is an example of a mixture. Tea is made up of milk, water, tea leaves and sugar all of which have different chemical compositions.

Mixture: It is a substance formed when two or more substances physically combine.

2.3 ALLOTROPES

The property of an element to exist in different physical forms is called allotropy. These different forms in the same physical state are called allotropes. Atoms of the same element arranged in different manners in the same physical state in allotropes. They are different structural forms of the same element. For example,

Diamond, graphite and buckyballs are three important allotropes of carbon.

Graphite:

Graphite is composed of flat two dimensional layers of hexagonally arranged carbon atoms. In a layer, each C-atom is covalently bonded to three other Carbon atoms. Weak intermolecular bonds exist between each layer which allows the layers to slide over one another without breaking the bonds. This arrangement makes graphite soft and slippery, making it ideal to be used as a lubricant. Graphite is a good conductor of electricity.

Diamond:

Diamond is the hardest and the purest crystalline allotrope of carbon. In its structure, each C-atom is covalently bonded to four other carbon atoms forming a rigid network of tetrahedral shape. This tetrahedral, three-dimensional arrangement makes it the hardest substance with a very high melting point. Since all the Carbon atoms are bonded with other carbon atoms, no free electrons are present resulting in the structure being non-conductive. Diamond is non-conductor of electricity.

Buckyballs (C-60):

Buckyballs, also known as fullerenes, have a football like fused hollow ring structure made up of twenty hexagons and twelve pentagons. Each of its 60 carbon atoms are bonded to 3 carbon atoms.

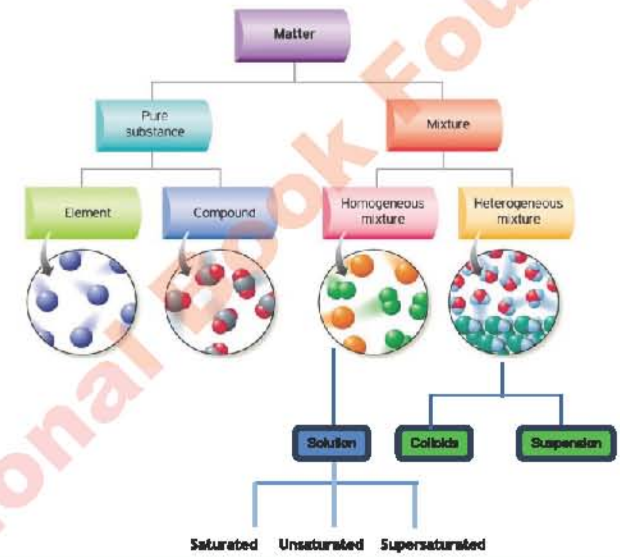


Fig 2.2: Classification of matter

2.4 Solution

A solution is a homogeneous mixture of two or more substances in which one substance is dissolved in the other. Homogeneous means that no particles or parts of different substances can be seen. When one substance dissolves, the solution looks exactly the same. **A substance that is dissolved is called a solute and a substance in which it is dissolved is called a solvent.** In solution, the particles are microscopic, less than 1 nm in diameter. A solution is a very stable mixture and the solute does not separate from the solvent itself.

Fig 2.3: Formation of Solution

In salt solution, salt is the solute and water is solvent. More than one solute may be present in a solution. For example, in soft drinks, water is a solvent while other substances like sugar, salts and CO_2 are solutes. Consider the example of air where Nitrogen gas is solvent and Oxygen, carbon dioxide and trace gases are solute. On the basis of physical states of solvent and solute can be categorized as solid, liquid and gaseous solutions. Generally, solutions are found in three physical states depending upon the physical state of the solvent, e.g. air is a gaseous, sea water is a liquid solution and alloy is a solid solution in real life.

Gaseous Solutions

In Gaseous Solutions solvent is a gas and solute can be a gas or liquid or solid. For example a mixture of nitrogen and hydrogen used in Haber's process (ammonia formation) and other is mixture of ammonia and carbon dioxide used for urea preparation. Fog, clouds and mist are examples of solutions where liquid water (solute) is dissolved in air (solvent). Smoke is a solution of carbon particle in gaseous air in our daily life.

Liquid Solutions

Carbonated drinks are solutions where solvent is liquid water and solute is gaseous carbon dioxide. Rectified spirit produced by fermentation of sugar cane, Vinegar (acetic acid in water), are examples of solutions where liquid dissolved in liquid. Brine and sugar syrup are solutions of solid salt and sugar in water.

Solid Solutions

Hydrogen gas on the nickel metal surface is used in ghee industry where hydrogen gas is solute and nickel catalyst is solvent. Solution of any metal (solid) in liquid mercury is called amalgam. Alloy industry is very common these days. Alloys are formed by mixing different metal (Brass, Bronze, steel).

2.4.1 Aqueous Solutions

Aqueous solution is formed by dissolving a substance in water. The dissolved substances in an aqueous solution may be solids, gases, or other liquids. In order to be a true solution, a mixture must be stable. For example, sugar in water and table salt in water. Water is called a universal solvent because it dissolves majority of compounds present in earth's crust. Aqueous solutions are mostly used in the laboratories.

Depending on amount of solute solution can be classified as

Solution

Saturated Solution | **Unsaturated Solution** | **Supersaturated solution**

2.4.2 Saturated Solution

A solution containing maximum amount of solute at a given temperature is called saturated solution.

When a small amount of solute at given temperature is added in a solvent, solute dissolves very easily in the solvent. If the addition of solute is kept on, a stage is reached when solvent cannot dissolve any more solute. At this stage, further added solute remains undissolved and it settles down at the bottom of the container. On the particle level, a saturated solution is the one, in which undissolved solute is in equilibrium with dissolved solute. At this stage, dynamic equilibrium is established. Although dissolution and crystallization continue at a given temperature, but the net amount of dissolved solute remains constant.

2.4.3 Unsaturated Solution

A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution. Such solutions have the capacity to dissolve more solute to become a saturated solution.

2.4.5 Supersaturated Solution

When saturated solutions are heated, they develop further capacity to dissolve more solute. Such solutions contain greater amount of solute than is required to form a saturated solution and they become more concentrated. The solution that is more concentrated than a saturated solution is known as supersaturated solution. Supersaturated solutions are not stable. Therefore, an easy way to get a supersaturated solution is to prepare a saturated solution at high temperature. It is then cooled to a temperature where excess solute crystallizes out and leaves behind a saturated solution.

Activity 2.1

Take 100g water in a beaker. Add a tea spoon of sugar in it. Stirr it. The sugar will dissolve.Repeat the process and the and added sugar will again dissolve in it. **A solution which can dissolve more of the solute at a given temperature is called an unsaturated solution.**

Go on adding sugar in the above solution till it starts settling down at the bottom of the beaker at a particular temperature. **The solution which cannot dissolve more solute at a particular temperature is called a saturated solution.**

Now heat the solution, stir it, add more sugar and it will dissolve. Go on adding more sugar and stir it. A stage will reach when no more sugar will dissolve and will start settling down at the bottom of the beaker. This solution is called supersaturated solution. A solution that contains more of the solute than is contained in the saturated solution is called supersaturated solution. How to know whether a solution is saturated or supersaturated? A supersaturated solution is not stable in the presence of crystals of solute. If you add a crystal of sodium thiosulphate to its saturated solution, it will simply drop to the bottom, without dissolving. But if you add a crystal of sodium thiosulphate to a supersaturated solution of sodium thiosulphate (see figure 2.4), crystallization will start. When crystallization has finished, you will have a saturated solution in the presence of sodium thiosulphate crystals.

Fig 2.4: Different types of solutions

2.4.6 Concentrated and Dilute Solution

The solutions are classified as dilute or concentrated on the basis of relative amount of solute present in them. Dilute solutions are those which contain relatively small amount of dissolved solute in the solution.

Concentrated solutions are those which contain relatively large amount of dissolved solute in the solution. For example, brine is a concentrated solution of common salt in water. These terms describe the concentration of the solution. Addition of more solvent will dilute the solution and its concentration decreases.

2.4.7 Solubility

Solubility is the maximum amount of solute which dissolves in a specified amount of solvent at a specific temperature. The solubility of a substance depends on the solvent used, as well as temperature and pressure. See Table 2.2.

2.4.8 Effect of Temperature on Solubility

The solubility of solutes depends on temperature. Depending on the nature of solute there is either:

a) **Increase in solubility with temperature e.g., KCl, NH_4Cl**

b) **Decrease in solubility with temperature e.g., Na_2SO_4 $Ca(OH)_2$**

Table 2.2: Solubility of some salts g/100g of solvent at Different Temperatures

(Solute)	Solubility (Amount of solute in 100g of solvent at 20°C	Solubility (Amount of solute in 100g of solvent at 100°C
NaCl	36.5g/100g H_2O	39.2g/100g H_2O
KCl	34.7g/100g H_2O	56g/100g H_2O
NH_4Cl	37.5g/100g H_2O	77g/100g H_2O
$Ca(OH)_2$	0.173g/100g H_2O	0.066g/100g H_2O

Example:

An example of a solute Whose decreases in solubility with increasing temperature is calcium hydroxide, which can be used to treat chemical burns and as an antacid.

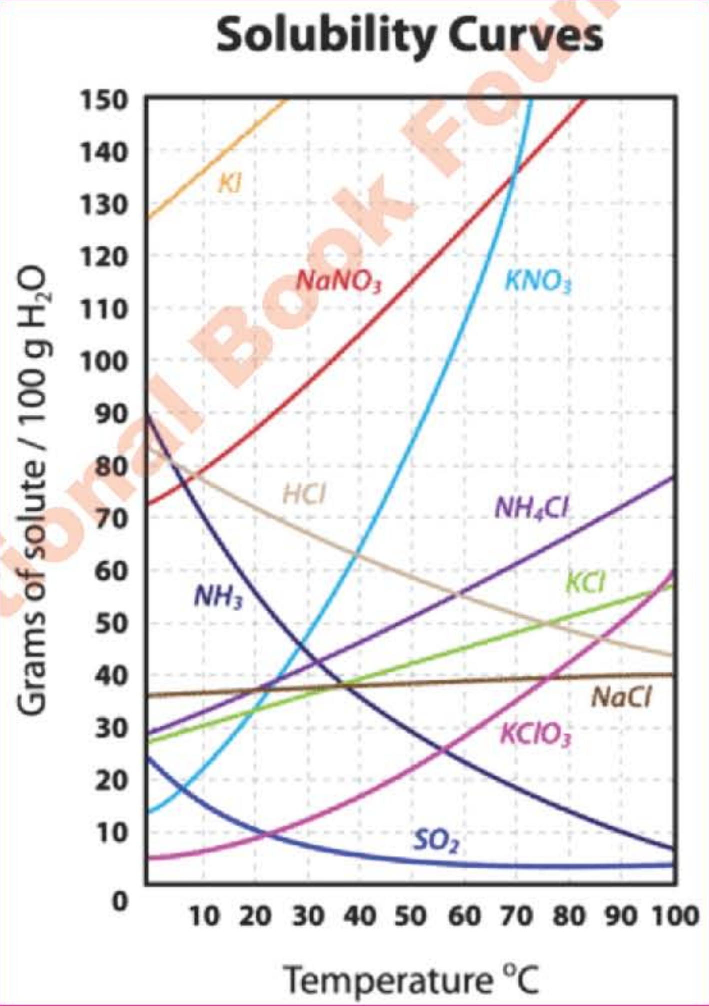


Fig.2.4 Effect of temperature on solubility

2.5 COLLOIDS & SUSPENSIONS

Colloid

These are heterogeneous mixtures in which the solute particles are larger than those present in the true solutions but not large enough to be seen by naked eye. A colloid is a mixture that has particles ranging between 1 and 1000 nanometers in diameter, yet are still able to remain evenly distributed throughout the solution. These are also known as colloidal dispersions because the substances remain dispersed and do not settle to the bottom of the container. The particles in such system dissolve and do not settle down for a long time. But particles of colloids are big enough to scatter the beam of light. It is called Tyndall effect. We can see the path of scattered light beam inside the colloidal solution. Tyndall effect is the main characteristic which distinguishes colloids from solutions. Hence, these solutions are called false solutions or colloidal solutions. Examples are starch, albumin, soap solutions, blood, milk, ink, jelly and toothpaste, etc.

Suspension

A suspension is defined as a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it. It is mixture of undissolved particles in a given medium. Particles are big enough (greater than 1000nm) to be seen with naked eyes. Examples are chalk in water (milky suspension), paints and milk of magnesia (suspension of magnesium oxide in water). For better understanding of true solutions, false solution and suspension, a comparison of their characteristics is given in table 2.3

Table 2.3

S.No	Solution	Colloids	Suspension
1	A homogeneous mixture of two or more components	A heterogeneous mixture of two or more components	A heterogeneous mixture of two or more components
2	Particle size vary from 0.1-1nm. Not visible by naked eye	Particle size vary from 1-10^3 nm. Visible by naked eye by naked eye	Particle size greater than 10^3 nm. Visible by naked eye by naked eye
3	Particles can pass through normal as well as ultra-filter paper	Particles can pass through normal filter paper but not through ultra-filter paper	Particles cannot pass through normal as well as ultra-filter paper
4	Cannot Scatter the light (due to small size)	Can Scatter the light (Tyndal effect)	Can Scatter the light (Tyndal effect)
5	Does not separate	Does not separate	Separate or settles down when stationary
Examples	Sea water	Milk	Muddy water

KEY POINTS

- Anything that has mass and occupies space is called matter.
- Plasma is an electrically charged gas, which is affected by electrical and magnetic fields.
- The property of an element to exist in different physical forms is called allotropy.
- Element: the simplest form of matter made up the same type of atoms
- Compound: A substance formed when two or more different atoms chemically combine.
- A homogeneous mixture of two or more components is called solution.
- Aqueous solution is formed by dissolving a substance in water.
- A solution containing maximum amount of solute at a given temperature is called saturated solution.
- A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution
- A colloid is a mixture that has particles ranging between 1 and 1000 nanometers in diameter
- A suspension is defined as a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it
- **References for additional information**
- Matter and its properties: Joseph Midthun, Paul Kobasa
- Cambridge IGSE™ Chemistry 5th Edition
- Cambridge International AS & A Level Chemistry (9701)

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Anything that has mass and occupies space is called.

(a) Liquid (b) Gas
(c) solid (d) Matter

(ii) Following are states of matter

(a) Gas (b) Liquid
(c) Solid (d) All of these

ii) Macroscopic properties are properties that can be visualized by

(a) the naked eye (b) microscope
(c) electron microscope (d) telescope

(iv) Matter can be described by both its

(a) physical properties and chemical properties.

(b) physical properties (c) chemical properties.

(v) A substance formed when two or more different combine chemically.

(a) atom (b) compound

(c) element (d) solution

2. **Give short answer.**

(i) Can you write the formula of the carbon dioxide gas that we exhale?

(ii) Define the element, Compound, Mixture

(iii) Differentiate between compound and mixture

(iv) Differentiate between concentrated and dilute solution

3. Define the term Allotropes Explain the allotropes of Carbon

4. What is difference between Homogeneous and heterogeneous solution?

5. Differentiate between the Colloids, Suspension

6. How can you identify solvent and solute?

7. If there are 18 protons in the Argon atom, then what is the atomic number of Argon?

8. Describe State of matter with example

9. Differentiate between the following.

a. Colloids and Suspensions

b. Elements and Compounds

c. Concentrated and Dilute solutions

10. Examine the concept of solubility.

THINK TANK

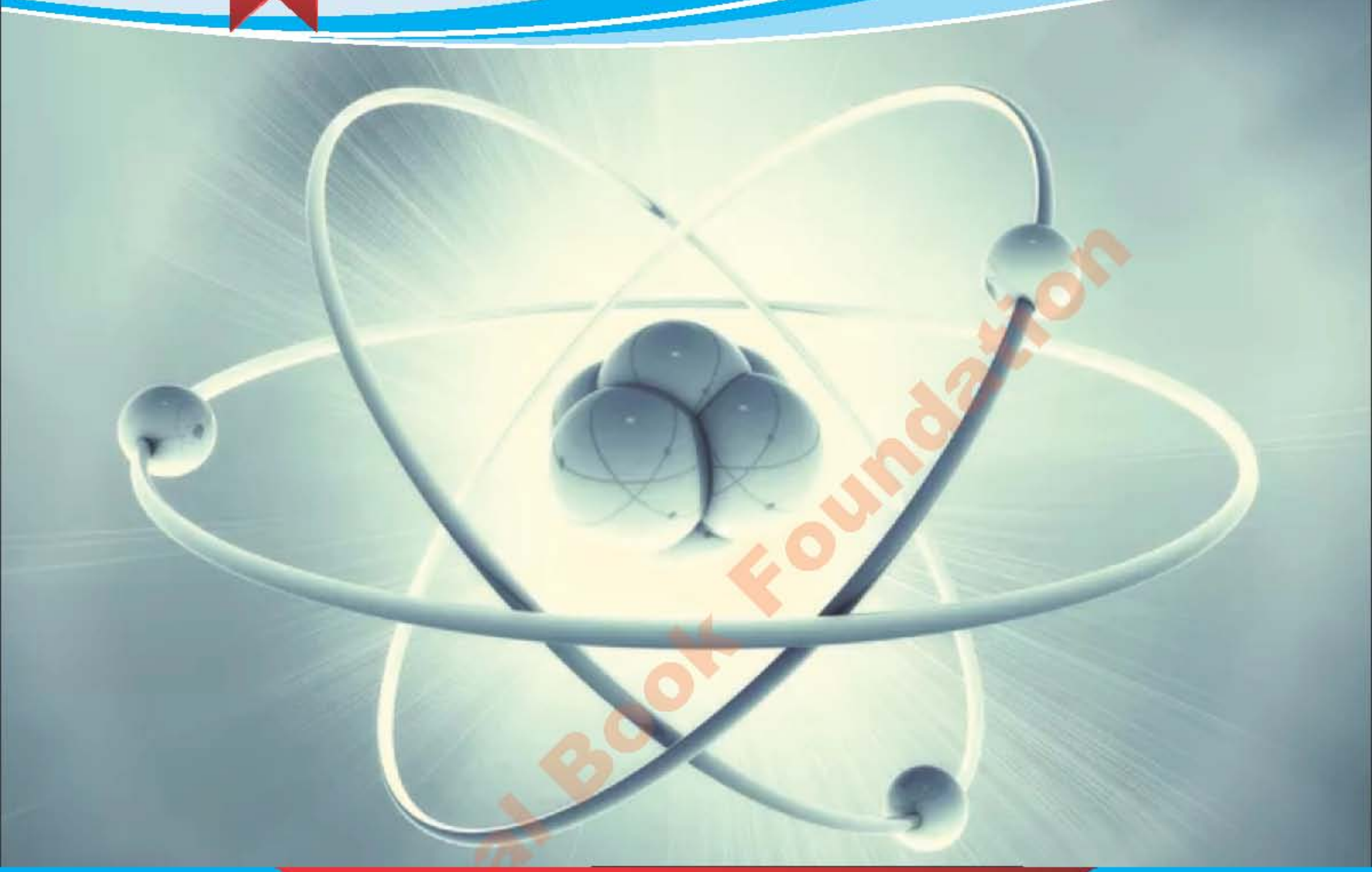
11. Why is a solution considered mixture?

12. How will you test weather given solution is a colloid or a solution?

PROJECT

Create a poster that illustrates the various form of matter in the students everyday environment.

UNIT 03



ATOMIC STRUCTURE

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Explain the structure of the atom as the central nucleus containing neutrons and protons surrounded by electrons in shells.
- State that, orbits(shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus.
- State the electrons are quantum particles with probabilistic paths whose exact paths and location cannot be mapped (with reference to uncertainty principle)
- Explain that nucleus is made up of protons and neutrons held together by strong nuclear force.
- Explain that an atomic model is an aid to understand the structure of an atom.
- State the relative charge and relative masses of a subatomic particles (an electron, proton, and neutron).
- Interpret the relationship between a subatomic particle, their mass, and charge.

- Illustrate the path that positively and negatively charged particles would take under the influence of a uniform electric field.
- Define proton number/atomic number as the number of protons in the nucleus of an atom.
- Explain that the proton number is unique to each element and use to arrange elements in periodic table.
- State that radioactivity can change the proton number and alter an atom`s identity.
- Define nucleon number/atomic mass as sum of protons and neutrons in the nucleus of an atom.
- Define isotopes as different atoms of the same element that have same number of protons but different neutrons.
- State that isotopes can affect molecular mass but not chemical properties of an atom.
- Determine the number of protons and neutrons of different isotopes.
- Define relative atomic mass as the average mass of isotopes of an element compared to1/12th of the mass of carbon-12
- State that isotopes can exhibit radioactivity.
- Discuss the importance of isotopes using carbon dating and medical imaging as examples.
- Describe the formation of positive(cation) and negative(anion) ions from atoms.
- Interpret and use the symbols for atoms and ions.
- Calculate the relative atomic mass from relative masses and abundance of isotopes.
- Calculate the relative mass of an isotope given relative atomic mass and abundance of all stable isotopes.

INTRODUCTION

This chapter presents the historical development of atomic theory to the modern atomic model. One of the basic concepts of atomic structure is atomic number and mass number, which define an element and its isotopes. Understanding the structure of atoms is essential to understanding many scientific phenomena.

3.1 ATOMIC MODELS

The concept of the atomic model evolved over time as our understanding of atomic structure deepened through experimental observations and theoretical advances. Several important models of the atom had been proposed throughout history, each contributing to the understanding of atomic behaviour and properties. The most important atomic models are:

Dalton's model

In 1803, the British chemist John Dalton presented a scientific theory on the existence and nature of matter. This theory is called Dalton's atomic theory. Main postulates of his theory are as follows:

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of a particular element are identical. They have same mass and same volume.
3. During chemical reaction atoms combine or separate or re-arrange. They combine in simple ratios.
4. Atoms can neither be created nor destroyed.

Dalton was able to explain quantitative results that scientists of his time had obtained in their experiments. He nicely explained the laws of chemical combinations. His brilliant work became the main stimulus for the rapid progress of the chemistry during nineteenth century. However, series of experiments that were performed in 1850's and beginning of twentieth century clearly demonstrated that atom is divisible and consists of subatomic particles, electrons, protons and neutrons.

In 1911 Rutherford performed an experiment in order to know the arrangement of electrons and protons in atoms.

Rutherford's Experiment

Rutherford bombarded a very thin gold foil about 0.00004cm thickness with α-particles. (figure 3.1). He used α-particles obtained from the disintegration of polonium. α-particles are helium nuclei that are doubly positively charged (He^{++}). Most of these particles passed straight through the foil. Only few particles were slightly deflected. But one in 1 million was deflected through an angle greater than 90° from their straight paths. Rutherford performed a series of experiments using thin foils of other elements. He observed similar results from these experiments.

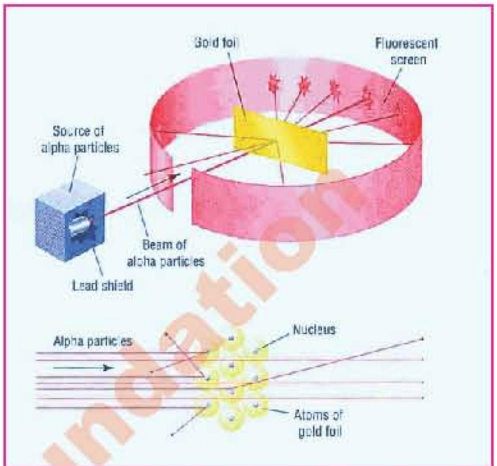


Figure 3.1 Rutherford's Experiment

Rutherford made the following conclusions:

1. Since majority of the α-particles passed through the foil undeflected, most of the space occupied by an atom must be empty.
2. The deflection of a few α-particles through angles greater than 90° shows that these particles are deflected by electrostatic repulsion between the positively charged α-particles and the positively charged part of atom.
3. Massive α-particles are not deflected by electrons.

On the basis of conclusions drawn from these experiments, Rutherford proposed a new model for an atom. He proposed a planetary model (similar to the solar system) for an atom. An atom is a neutral particle. The mass of an atom is concentrated in a very small dense positively charged region. He named this region as nucleus. The electrons are revolving around the nucleus in circles. These circles are called orbits. The centrifugal force due to the revolution of electrons balances the electrostatic force of attraction between the nucleus and the electrons.

Defects in Rutherford's Atomic Model

Rutherford's model of an atom resembles our solar system. It has following defects:

1. Classical physics suggests that electron being charged particle will emit energy continuously while revolving around the nucleus. Thus the orbit of the revolving electron becomes smaller and smaller until it would fall into the nucleus. This would collapse the atomic structure.
2. If revolving electron emits energy continuously it should form a continuous spectrum.

Bohr's Atomic Theory

In1913 Neil Bohr, proposed a model for an atom that was consistent with Rutherford's model. But it also explains the observed line spectrum of the hydrogen atom. Main postulates of Bohr's atomic theory are as follows:

1. The electron in an atom revolves around the nucleus in one of the circular orbits. Each orbit has a fixed energy. So each orbit is also called energy level.
2. The energy of the electron in an orbit is proportional to its distance from the nucleus. The farther the electron is from the nucleus, the more energy it has.
3. The electron revolves only in those orbits for which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$ where h is Plank's constant (its value is 6.626×10^{-34} J.s).
4. Light is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit. Electron present in a particular orbit does not radiate energy.
5. The energy of the light emitted is exactly equal to the difference between the energies of the orbits.

$$\Delta E = E_2 - E_1$$

Where ΔE is the energy difference between any two orbits with energies E_1 and E_2

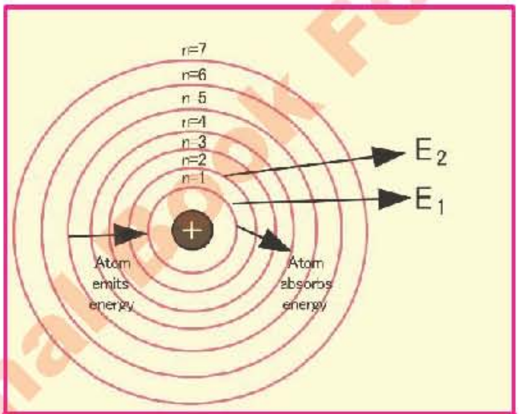


Figure 3.2 Bohr's model of the atom

Bohr model does not depict the three dimensional aspect of an atom.

Quantum Mechanical Model:

This is the current model used by modern science to describe the structure of the atom. It incorporates the principles of quantum mechanics and treats electrons as wave-particle entities. Instead of exact orbits, it defines probability regions, called orbitals, where electrons are likely to be found.

The Heisenberg Uncertainty Principle :

Heisenberg uncertainty principle is one of the fundamental concepts of quantum mechanics and is named after the German physicist Werner Heisenberg, who formulated it in 1927.

This principle states that it is impossible to simultaneously determine the exact location and future trajectory of an electron. As a result, plotting the electron orbit around the nucleus becomes an irresistible challenge.

Imagine that you have a single hydrogen atom and you decide to observe the position of that single electron at a given moment. Shortly after you repeated this process, the electron moved to another position. This means that from the original location to the next one is completely unknown to you. Continuous repetition of this process allows the gradual construction of a three-dimensional map representing the likely locations where the electron is expected to exist. You cannot know for sure where an electron is and where it goes next. This makes it impossible to draw the orbit of the electron around the nucleus.

In hydrogen, the electron has the potential to exist anywhere in the spherical region surrounding the nucleus. 95% (or whatever you want) of the time, the electron will be in a relatively simple region of space close to the nucleus, called an orbital. An orbital is the region of space where the electron lives.

Louis de Broglie, a French physicist, in 1924 proposed duel nature of electrons. He suggested that sub-atomic particles like electrons, can exhibit both particle-like and wave-like behaviour. His idea opened the door for new possibilities in understanding behaviour of sub-atomic particles. This concept made a significant contribution to the development of quantum mechanics.

In 1927, Davisson and Germer, experimently confirmed the de Broglie hypothesis that electron has wave like behaviour. This discovery laid the foundation for the Modern Quantum Mechanics.

Understanding Atomic Models

An atomic model is a tool for understanding the structure and behavior of atoms and their interactions in chemical reactions. Any atomic model helps us understand the structure of an atom. An atomic model is not a physical model, but represents a conceptual imagination. This helps to explain experimental observations of atomic behavior. The atomic model gives us a simplified representation of complex reality. As research and technology progress, scientists continue to improve our knowledge and atomic models.

A simple view of the structure of an atom

The nucleus of an atom is in the center. It contains protons and neutrons. Protons and neutrons are collectively called nucleons. The nucleus is surrounded by electrons in shells. Protons and neutrons are massive particles. The mass of electrons is so small. So, in practice, the mass of an atom is concentrated in the nucleus.

Nuclear Force

The nucleus contains protons and neutrons. Protons are positively charged and neutrons are neutral. The nucleus has no negative charge. The positively charged protons must cancel each other out and the nucleus must break apart. But atoms are stable and have existed for billions of years. Therefore, there must be some kind of attraction that connects them. No electrostatic or magnetic forces occur within the core. This is because these forces involve both attraction and repulsion. Therefore, the force that binds protons and neutrons together is a strong force. This force is called strong nuclear force. This is defined as the strong attractive force that binds protons and neutrons together. This force is stronger than electrostatic or magnetic forces. This force exists between neutrons and neutrons, protons and protons, and neutrons and protons.

3.2 SUBATOMIC PARTICLES

Subatomic particles are the fundamental particles that make up atoms. The three main subatomic particles are:

Proton:

- Relative charge: +1
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6726×10^{-27} kg

Neutron:

- Relative charge: 0 (neutral)
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6749×10^{-27} kg

Electron:

- Relative charge: -1
- Relative mass: Approximately 1/1836 amu or 9.11×10^{-31} kg

Protons and neutrons are found in the nucleus of an atom, whereas electrons orbit around the nucleus in energy levels or shells. They play crucial roles in determining the properties and behaviour of atoms and molecules. Neutrons and protons are held together in the nucleus by a strong nuclear force. This force exists between neutron-neutron, proton-proton, and neutron-proton.

Relationships between subatomic particles:

Protons and neutrons have roughly the same mass, around 1 amu. This mass contributes significantly to the total mass of the atom. Electrons have much less mass, so their contribution to the total mass of an atom is usually negligible.

The interaction between the negatively charged electrons and positively charged protons in the nucleus is what holds the atoms together.

The behavior of protons, neutrons and electrons in an electric field

What happens when a beam of these particles passes between two electrically charged plates?

figure

- Protons are positively charged and are deflected on a curved path toward the negative plate.
- Electrons are negatively charged and are deflected on a curved path toward the positive plate.
- Neutrons have no charge, go straight ahead.
- If the electrons and protons are traveling at the same speed, the electrons being lighter are deflected far more strongly than the heavier protons.

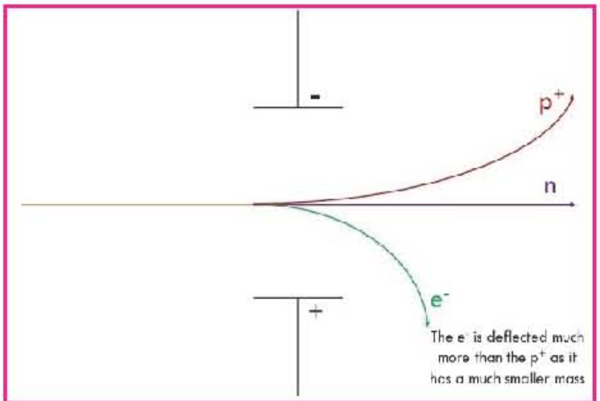


Figure 3.3: Path of positively and negatively charged particles through the uniform electric field.

Charge Neutrality

Atoms are electrically neutral because the number of protons (positively charged) in the nucleus is equal to the number of electrons (negatively charged) in the electron cloud. The charges balance each other so there is no net charge on the atoms.

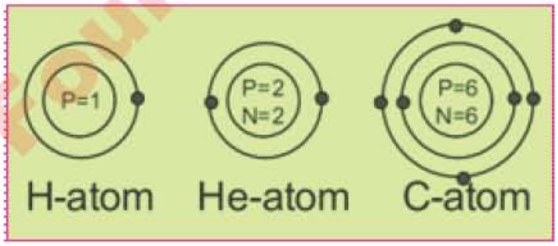
Radioisotopes

Different isotopes of the same element have the same number of protons in their atomic nuclei but differing numbers of neutrons. Some isotopes of an element are unstable and show radioactive decay. Radioactive isotopes of an element can be defined as atoms that contain an unstable combination of neutrons and protons, or excess energy in their nucleus. For example, hydrogen-3(protium), carbon-14, uranium-238 etc.

3.3 PROTON OR ATOMIC NUMBER

What determines the identity of an element?

Proton number refers to the number of protons in the nucleus of an atom. It is also known as the atomic number and is indicated by the symbol "Z".



Protons have a positive electrical charge. In neutral atoms, the number of protons is equal to the number of electrons. This balances the positive charge of the protons. This means that the proton number also indicates the number of electrons in the atom. For example, there is only one proton in the nucleus of a H atom; therefore its atomic number is 1. All the atoms of a given element have the same number of protons and therefore the same atomic number.

Do you think atomic number of He is 2? What is the proton number of C-atom?

Uniqueness of proton number

Each element has a unique proton number that distinguishes it from other elements. It determines the various properties of an element and its position in the periodic table. In the periodic table, elements are arranged based on their atomic or proton number. Therefore, the number of protons is related to the position of the element in the periodic table. Thus, the number of protons determines each particular element. This will tell you what element you are talking about.

For example, if an atom has a proton number of 6, it must be carbon. If an atom has 11 protons, it must be sodium. Similarly, each nitrogen atom has 7 protons, each oxygen atom has 8 protons, etc. You can identify each atom by the number of protons.

Nucleon number or Atomic mass

The total number of protons and neutrons in an atom is known as its mass number or nucleon number.

Some atoms of an element have different number of neutrons, such atoms are called isotopes.

No. of neutrons = mass number – atomic number

Example 3.1: Determining the number of protons and neutrons in an atom

Atomic number of an element is 17 and mass number is 35. How many protons and neutrons are in the nucleus of an atom of this element?

Problem Solving Strategy:

Number of protons are equal to atomic number and

Number of neutrons = mass number - atomic number

Solution:

Number of protons = atomic number= 17

Number of neutrons = mass number - atomic number

= 35-17= 18

Radioactivity

The proton number determines the identity of the element. In stable elements, the nuclear force is balanced. In some elements, the nuclear forces are not naturally balanced. The nucleus of these atoms decays and becomes another atom. This process is called radioactive decay and this phenomenon is called radioactivity. This process continues until the forces in the nuclear core are balanced. In radioactive decay, when an atom emits a neutron, it changes to another isotope of that atom. But when it emits a proton, it becomes another atom. This means that radioactivity can change the number of protons in an atom and thus change the identity of the atom. For example;

1. Carbon-14 is a radioactive isotope of carbon. It is naturally present in the atmosphere. When any living organism takes in carbon dioxide from the air, it incorporates both C-14 and C-12 atoms into its tissues. The radioactive C-14 undergoes radioactive decay, transforms into nitrogen-14.
2. Uranium-238 is a radioactive isotope of uranium. It decays over time and finally transforms into stable lead-206 atom.

3.4 RELATIVE ATOMIC MASS AND ATOMIC MASS UNIT

The first quantitative information about atomic masses came from the work of Dalton, Gay Lussac, Lavoisier, Avogadro and Berzelius. By observing the proportions in which elements combine to form various compounds, nineteenth century chemists calculated relative atomic masses. An atom is extremely small particle, therefore, we cannot determine the mass of a single atom. However, it is possible to determine the mass of one atom of an element relative to another experimentally. This can be done by assigning a value to the mass of one atom of a given element, so that it can be used as standard. By international agreement in 1961, light isotope of carbon C-12 has been chosen as a standard. This isotope of carbon(C-12) has been assigned a mass of exactly 12 atomic mass unit. This value has been determined accurately using mass spectrometer. The mass of atoms of all other elements are compared to the mass C-12. Thus **"the mass of an atom of an element relative to the mass of an atom of C-12 is called its relative atomic mass".**

One atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one C-12 atom.

$$\text{Mass of one C-12 atom} = 12 \text{ amu}$$

$$1\text{amu} = \frac{\text{mass of one C-12 atom}}{12}$$

A hydrogen atom is 8.40% as massive as the standard C-12 atom. Therefore, relative atomic mass of hydrogen.

$$= \frac{8.40}{100} \times 12 \text{ amu}$$

$$= 1.008 \text{ amu}$$

Similarly, relative atomic masses of O, Na, Al are 15.9994 amu, 22.9898 amu, 26.9815 amu respectively. Table 3.1 shows the relative atomic masses of some elements.

Table 3.1: relative atomic masses of some elements

Element	Relative atomic mass	Element	Relative atomic mass
H	1.008 amu	Al	26.9815 amu
N	14.0067amu	S	32.06 amu
O	15.9994amu	Cl	35.453 amu
Na	22.9898 amu	Fe	55.847 amu

3.5 ISOTOPES

Figure 3.1 shows Bohr's Model for two atoms A and B. Can you identify three similarities and two differences in these atoms?

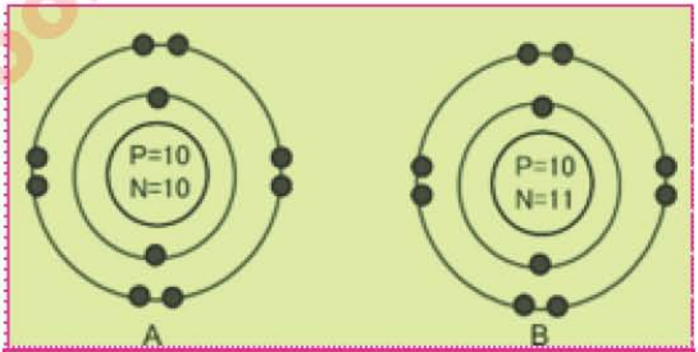


Figure 3.1

You will find,

(a) Both the atoms have same number of protons.

(b) Both the atoms have same number of electrons.

(c) Both have same atomic number.

(d) Both have different number of neutrons.

(e) Both differ in total number of protons and neutron. This means they have different mass numbers.

Since both the atoms have the same atomic number, they must be the atoms of same element and are called isotopes. The word isotope was first used by Soddy. It is a Greek word "isos" means same and "tope" means place.

Isotopes are atoms of an element whose nuclei have the same atomic number but different mass number. This is because atoms of an element can differ in the number of neutrons. Isotopes are chemically alike and differ in their physical properties.

How does the discovery of isotopes contradicted Dalton's atomic theory?

3.5.1 Isotopes of Hydrogen

Hydrogen has three isotopes. Hydrogen -1 (Protium) has no neutron. Almost all the hydrogen is Hydrogen -1. Its symbol is $^{1}_{1}H$. Hydrogen - 2 (deuterium) has one neutron and hydrogen -3 (Tritium) has two neutrons. Their symbols are $^{2}_{1}H$ and $^{3}_{1}H$ respectively Because hydrogen -1 also known as protium has only one proton, adding a neutron doubles it mass.

Protium / Hydrogen is a colourless, odourless, and tasteless gas. It is insoluble in water and is highly inflammable gas. Water that contain hydrogen-2 atoms in place of hydrogen-1 is called heavy water. Table 3.2 Shows some physical properties of ordinary water and heavy water.

Table 3.2 - Comparision of ordinary water and heavy water.

Property	Ordinary water	Heavy water
Melting Point	$0.00^{0}C$	$3.81^{0}C$
Bioling point	$100^{0}C$	$101.2^{0}C$
Density at $25^{0}C$	0.99701 g/cm^3	1.1044 g/cm^3

Isotopes affect molecular mass of a substance, can change physical properties but do not change chemical properties.

At what temperature would a sample of heavy water freeze?

Naturally occurring hydrogen contains 99.99% protium, 0.0015% Deuterium. Tritium is radioactive and is rare. Tritium is not found in naturally occurring hydrogen because its nucleus is highly unstable.

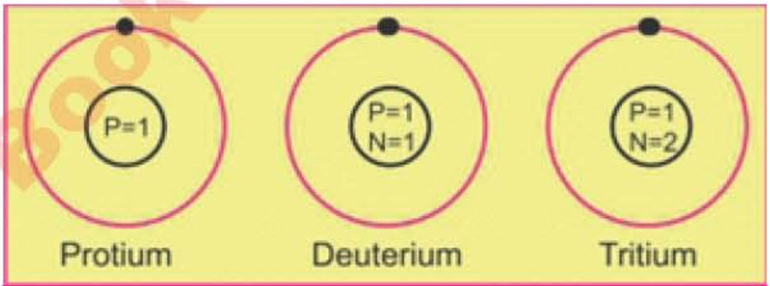


Figure 3.2 Isotopes of hydrogen

3.5.2 Isotopes of Carbon

Carbon has three isotopes. Carbon-12, carbon-13 and carbon -14. Almost all the carbon is carbon-12. Its symbol is. $^{12}_{6}C$ It has six neutrons and six protons. Carbon-13 has symbol. $^{13}_{6}C$ It has seven neutrons and six protons. Carbon-14 has eight neutrons and six protons. Its symbol is $^{14}_{6}C$ Different forms of carbon are black or greyish black solids except diamond. They are odourless and tasteless. They have high melting and boiling points and are insoluble in water.

Activity 3.1

Carbon has three isotopes $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$ Figure 3.3 shows incomplete structure of isotopes of carbon. Can you complete it?

Natural abundance of isotopes of carbon is as follows

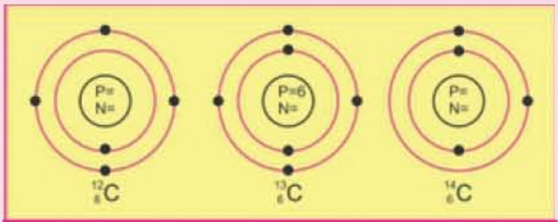


Figure 3.3 Isotopes of Carbon

$$^{12}_{6}C = 98.8\%, \quad ^{13}_{6}C = 1.1\%, \quad ^{14}_{6}C = 0.009\%$$

3.5.3 Isotopes of Chlorine

There are two natural isotopes of chlorine, chlorine-35 and chlorine-37. An atom of chlorine-35 has 17 protons and 18 neutrons. An atom of chlorine-37 has 17 protons and 20 neutrons. Chlorine-35 occurs in nature about 75% and chlorine-37 about 25%. Chlorine is a greyish yellow gas with sharp pungent irritating smell. It is fairly soluble in water

Activity 3.2

Chlorine has two isotopes. Figure 3.4 shows the structure of isotopes of chlorine. Can you write isotope symbol for each?

Isotope symbols:	______	______
Natural abundance	75.77%	24.23%

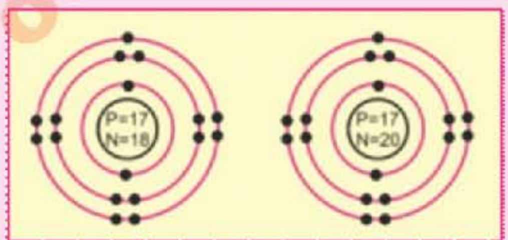


Figure 3.4 Isotopes of chlorine

3.5.4 Isotopes of Uranium

Activity 3.3

Uranium has three isotopes with mass number 234, 235 and 238 respectively.

$$^{234}_{92}U, \quad ^{235}_{92}U, \quad ^{238}_{92}U$$

The $^{235}_{92}U$ isotope is used in nuclear reactors and atomic bombs, whereas the $^{238}_{92}U$ isotope lacks the properties necessary for these applications. $^{234}_{92}U$ is rare. Natural abundance of Uranium isotopes is as follows

$$^{234}_{92}U = 0.006\%, \ ^{235}_{92}U = 0.72\%, \ ^{238}_{92}U = 99.27\%$$

Fill in the blanks?

$^{234}_{92}U$has ___ protons, ___ electrons and ___ neutrons

$^{235}_{92}U$has ___ protons, ___ electrons and ___ neutrons

$^{238}_{92}U$ has ___ protons, ___ electrons and ___ neutrons

When uranium-238 decays into thorium-234, it emits alpha particle. An alpha particle is doubly positively charged helium nucleus.

$$^{238}_{92}U \longrightarrow ^{234}_{90}Th + ^{4}_{2}He$$

The fission of uranium-235 yields smaller nuclei, neutron and energy. The nuclear energy released by the fission of one kilogram of uranium-235 is equivalent to chemical energy produced by burning more than 17000 kg of coal.

Chemical properties of an element depend upon the number of protons and electrons. Neutrons do not take part in ordinary chemical reactions. Therefore, isotopes of an element have similar chemical properties.

3.5.5 Determination of Relative Atomic Mass

The relative atomic mass of an element can be calculated from the relative masses of its isotopes and their relative abundance.

Natural abundance of isotopes of carbon is as follows

$^{12}_{6}C = 98.8\%$, $^{13}_{6}C = 1.1\%$, $^{14}_{6}C = 0.009\%$

Calculate relative atomic mass of carbon.

Solution:

The relative atomic mass is a weighed average of the all the naturally occurring isotopes of an element, taking into consideration of their natural abundance. Use general formula

$$\text{Relative atomic mass of C} = \frac{\text{RA of C-12} \times \text{at.mass of C-12} + \text{RA of C-13} \times \text{at.mass of C-13} + \text{RA of C-14} \times \text{at.mass of C-14}}{100}$$

$$\text{Relative atomic mass of C} = \frac{98.8 \times 12 + 1.1 \times 13 + 0.009 \times 14}{100}$$

$$\text{Relative atomic mass of C} = \frac{1185.6 + 14.3 + 0.126}{100}$$

$$\text{Relative atomic mass of C} = 12.00026\,\text{amu}$$

CONCEPT ASSESSMENT EXERCISE 3.1

An element has two isotopes A and B.

The relative atomic mass of element is 35.5 amu. Relative abundance of isotope A is 75.77 % and its isotopic mass is 35. Find the isotopic mass of B if its relative abundance is 24.23 %.

3.5.6 Uses of Isotopes

Stable and radioactive isotopes have many applications in science and medicines. Some of these are as follows:

(i) Radioactive iodine -131 is used as a tracer in diagnosing thyroid problem.

(ii) Na-24 is used to trace the flow of blood and detect possible constrictions or obstructions in the circulatory system.

(iii) Iodine-123 is used to image the brain.

(iv) Cobalt-60 is commonly used to irradiate cancer cells in the hope of killing or shrinking the tumors.

(v) Carbon-14 is used to trace the path of carbon in photosynthesis. Radioactive isotopes are used to determine the molecular structure e.g. sulphur-35 has been used in the structure determination of thiosulphate, $S_2O_3^{-2}$ ion.

(vi) Radioactive isotopes are also used to study the mechanism of chemical reactions.

(vii) Radioactive isotopes are used to date rocks, soils, archaeological objects, and mummies.

3.5.7 Carbon Dating

Carbon-14 is used to estimate the age of carbon-containing substances. Carbon atoms circulate between the oceans, and living organism at a rate very much faster than they decay. As a result the concentration of C-14 in all living things, keep on increasing. After death organisms no longer pick up C-14. By comparing the activity of a sample of skull or jaw bones, with the activity of living tissues, we can estimate how long it has been since the organism died. This process is called dating.

3.6 CATIONS AND ANIONS:

Cations:

Cations are positively charged ions that form when an atom loses one or more electrons. Cations are usually formed from metal atoms that tend to lose electrons to achieve a stable electron configuration similar to a noble gas. When an atom loses one or more electorns, it forms a cation. The resulting cation has the electronic configuration of a noble gas. Neutral atoms have equal number of protons and electrons. When an atom loses one or more electrons, the number of protons becomes greater than electrons, as a result atom acquires positive charge.

Example 3.1: Describing the formation of cations

Describe the formation of Na^+ and Mg^{+2} cations.

Problem Solving Strategy:

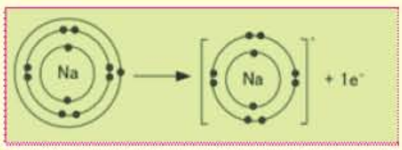
1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. Sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group IIA in the periodic table. It has two valence electrons. Magnesium atom loses these electrons to achieve noble gas configuration. Represent this by drawing the complete electronic configuration or using an electron dot structure. This number also corresponds to the Group number in the periodic table.

Solution:

(a) Formation of Na^+ ion

$$Na\ 1s^2 2s^2 2p^6 3s^1 \xrightarrow{-e^-} Na^+ 1s^2 2s^2 2p^6$$

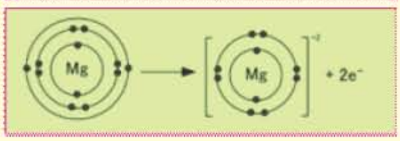
You can also represent this by following electron dot structure,



(b) Formation of Mg^{+2} ion

$$Mg\ 1s^2 2s^2 2p^6 3s^2 \xrightarrow{-2e^-} Mg^{+2}\ \ 1s^2 2s^2 2p^6$$

You can also represent this by electron dot structure,



CONCEPT ASSESSMENT EXERCISE 3.2

Describe the formation of cations for the following metal atoms:

(a) Li (atomic no 3)

(b) Al (atomic no.13)

Anions

Anions are negatively charged ions that form when an atom gains one or more electrons. This process usually occurs when an atom has a relatively high electron affinity, meaning that it can easily attract and capture more electrons to achieve a stable electron configuration similar to a noble gas. When an atom gains one or more electrons, the number of electrons becomes greater than protons, so it acquires negative charge.

Example 3.1: Describing the formation of anions.

Describe the formation of anions for the following non-metal atoms:

(a) Oxygen(atomic no.8) (b) Fluorine (atomic no. 9)

Problem Solving Strategy:

1. Write electronic configuration or dot structure.
2. Find the number of electrons needed to acquire eight electron configuration.
3. Represent addition of electrons.

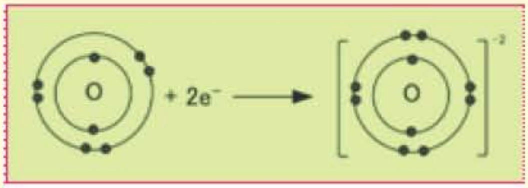
Solution:

(a) Formation of anion by oxygen atom.

Oxygen belongs to Group VIA on the periodic table. So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.

$$O\ 1s^2 2s^2 2p^4 + 2e^- \longrightarrow O^{-2}\ 1s^2 \underbrace{2s^2 2p^6}_{\text{octet}}$$

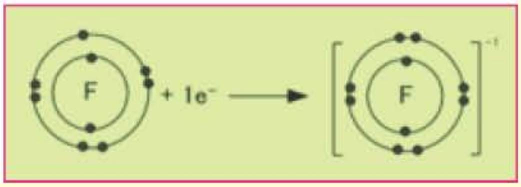
You can also represent this by electron dot structure,



(b) Formation of anion by fluorine atom

Fluorine belongs to Group VIIA on the periodic table. So it has seven electrons in the valence shell. A fluorine atom therefore, requires only one electron to complete octet.

$$F\ 1s^2 2s^2 2p^5 + e^- \longrightarrow F^-\ 1s^2 \underbrace{2s^2 2p^6}_{\text{octet}}$$



You can also represent this by electron dot structure,

CONCEPT ASSESSMENT EXERCISE 3.3

Describe the formation of anions by the following non-metals.

(a) Sulphur (atomic No. 16)

(b) Chlorine(atomic No. 17)

3.7 ELECTRONIC CONFIGURATION

To understand electronic configuration, you should know about shells and sub-shells.

Shells

According to Bohr's atomic theory, the electron in an atom revolves around the nucleus in one of the circular paths called shells or orbits. Each shell has a fixed energy. So each shell is also called energy level. Each shell is described by an n value. n can have values 1,2,3.....

When,

n = 1, it is K shell

n = 2, it is L shell

n = 3, it is M shell etc.

As the value of n increases the distance of electron from the nucleus and energy of the shell increases.

Sub-Shells

A shell or energy level is sub divided into sub-shells or sub-energy levels. n value of a shell is placed before the symbol for a sub-shell. For instance,

n = 1, for K shell. It has only one sub-shell which as represented by 1s.

For L shell n =2, L shell has two sub-shells, these are designated as 2s and 2p.

For M shell n =3 So M shell has 3 sub-shells called 3s , 3p and 3d. While N shell has 4s, 4p, 4d and 4f sub-shells.

s sub-shell can accommodate maximum 2 electrons.

p sub-shell can accommodate maximum 6 electrons.

d sub-shell can accommodate maximum 10 electrons.

f sub-shell can accommodate maximum 14 electrons.

The increasing order of energy of the sub-shells belonging to different shells is given below.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \ldots..$$

The arrangement of electrons in sub-shells is called as the electronic configuration. We can fill the electrons present in various elements by using Auf Bau Principle. According to this principle, electrons fill the lowest energy sub-shell that is available first. This means electron will fill first 1s, then 2s, then 2p and so on.

Symbols for atoms and ions

The symbol for an atom represent the element . It consists of one or two-letters, the mass number as a left superscript, the atomic number as a left subscript, and the charge as a right superscript. For example;

$^{24}_{12}Mg^{2+}$

This number is often omitted. This diagram shows symbol for magnesium "Mg" which stands for magnesium. The number to the upper left of the symbol is the mass number, which is 24. The number to the upper right of the symbol is the charge which is positive 2. The number to the lower left of the symbol is the atomic number which is 12.

KEY POINTS

- Rutherford proposed a planetary model for an atom. The nucleus of an atom is composed of protons. The electrons surround the nucleus and occupy most of the volume of the atom.
- According to Bohr's atomic model, the electron in an atom revolves around the nucleus in fixed circular orbits called shells. Isotopes are atoms of an element that differ in the number of neutrons.
- $^{235}_{92}U$ isotope is used in nuclear reactors and atomic bombs.
- Radioactive isotopes have many applications in science and medicines such as killing cancer cells, diagnosing thyroid problem, to image the brain, to detect obstruction in the circulatory system, to date rocks, soils, mummies etc.
- A shell or energy level is divided into sub-shells.
- There are four types sub-shell s, p, d, and f.
- The arrangement of electrons in sub-shells is called as the electronic configuration.
- According to the Auf Bau Principle, electrons fill the lowest energy levels first.
- **References for additional information**
- B.Earl and LDR Wilford, Introducion to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Chlorine has two isotopes, both of which have
(a) same mass number. (b) same number of neutrons.
(c) different number of protons. (d) same number of electrons.

(ii) Number of neutrons in $^{27}_{13}M$ are
(a) 13 (b) 14
(c) 27 (d) 15

(iii) Which isotope is commonly used to irradiate cancer cells?
(a) Iodine-123 (b) Carbon-14
(c) Cobalt-60 (d) Iodine-131

(iv) M shell has sub-shells:
(a) 1s, 2s (b) 2s, 2p
(c) 3s, 3p, 3d (d) 1s, 2s, 3s

(v) A sub-shell that can accommodate 6 electrons is
(a) s (b) d
(c) p (d) f

(vi) $_{11}Na$ has electronic configuration:
(a) $1s^2,2s^2,3s^1$ (b) $1s^2,2s^2,2p^7$
© $1s^2,2s^2,2p^5,3s^2$ (d) $1s^2,2s^2,2p^6,3s^1$

(vii) Which of the following statement is not correct about isotopes?
(a) they have same atomic number
(b) they have same number of protons
(c) they have same chemical properties
(d) they have same physical properties

(viii) Which isotope is used in nuclear reactors?
(a) U-234 (b) U-238
(c) U-235 (d) All of these

2. Give short answer.

(i) Distinguish between shell and sub-shell
(ii) Why an atom is electrically neutral?
(iii) How many sub-shells are there in N shell.
(iv) Give notation for sub-shells of M shell.
(v) List the sub-shells of M Shell in order of increasing energy
(vi) Can you identify an atom without knowing number of neutrons in it?

3. The electronic configurations listed are incorrect. Explain what mistake have been made in each and write correct electronic configurations.

4. $x = 1s^2, 2s^2, 2p^4, 3p^2$, $y = 1s^2, 2s^1, 2p^1$, $z = 1s^2, 2s^2, 2p^5, 3s^1$

5. Which orbital in each of the following pairs is lower in energy?

6. (a) 2s, 2p (b) 3p, 2p (c) 3s, 4s

7. Draw Bohr's Model for the following atoms indicating the location for electron, protons and neutrons:

8. (a) Potassium (Atomic No 19, Mass No. 39)

   (b) Silicon (Atomic No. 14 Mass No. 28)

   (c) Argon (Atomic No. 18 Mass No. 39)

9. Write electronic configuration for the following elements:

10. (a) $^{28}_{14}Si$ (b) $^{24}_{12}Mg$ © $^{27}_{13}Al$ (d) $^{40}_{18}Ar$

11. State the importance and uses of isotopes in various fields of life.

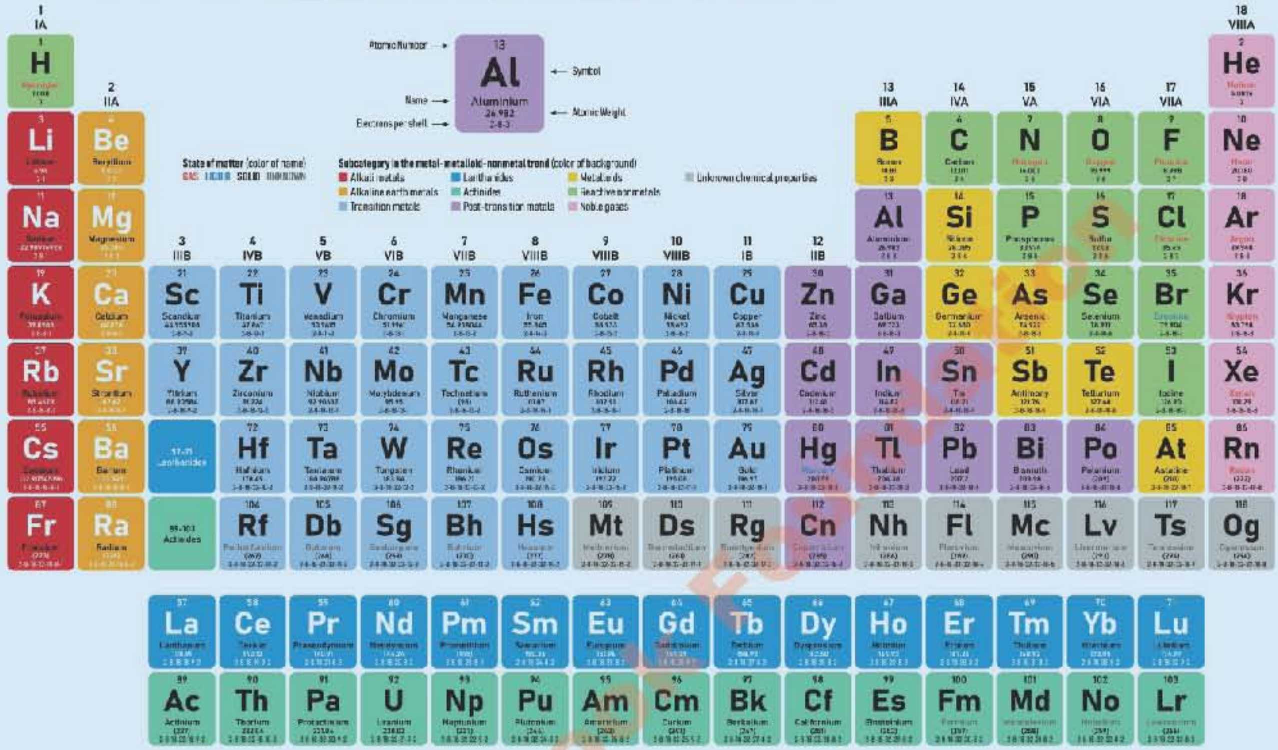
12. The atomic number of an element is 23 and its mass number is 56.

    a. How many protons and electrons does an atom of this element have?

    b. How many neutrons does this atom have?

13. The atomic symbol of aluminium is written as $^{27}_{13}Al$. What information do you get from it?

UNIT 04



PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define the periodic table as an arrangement element in periods and group, in order of increasing proton number/atomic number.
- Identify the group or period or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced).
- Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells.
- Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration.
- Identify trends in groups and periods, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity, and density.
- Use terms like alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides, and actinides in reference to the periodic table.

- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity.
- Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties.
- Define Group 1 Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.
- Predict properties of other elements in Group 1, given information about the elements.
- Predict properties of elements in Group 1 in order of reactivity given relevant information.
- Define Group VII halogens as diatomic non-metals with general trends limited to increasing density and decreasing reactivity.
- Identify the appearance of halogens at rtp as fluorine as pale yellow gas, chlorine as yellow-green gas, bromine as red-brown liquid, iodine as grey-black solid.
- Explain the displacement reactions of halogens with other halide ions and also as reducing agents.
- Predict the properties of elements in group VII, given information about the elements.
- Analyze the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths.
- Describe the transition elements as metals that: have high densities, high melting points, variable oxidation numbers, form coloured compounds and act as catalysts for industrial purposes. (some example include catalysts being used are the Haber process, catalytic converters, Contact process and manufacturing of margarine).
- Define the Group 18 noble gases as un-reactive mono-atomic gases.
- Explain this interms of electronic configuration.
- Compare the general physical properties of metals and non-metals. (specifically in terms of
  - a. Thermal conductivity
  - b. Electrical conductivity
  - c. Malleability and ductility
  - d. Melting points and boiling points

INTRODUCTION

Welcome to the exciting world of chemistry, where the elements come to life thanks to the remarkable periodic table. From its humble beginnings, where only 23 elements were known until the end of the 18th century, to its development of 118 elements today. It is very difficult and impossible to remember information about the reactions, properties, and atomic masses of elements. So we obviously need a way to organize our information about them. The periodic table is one of the most important tools in chemistry. It is very useful for understanding and predicting the properties of elements. For example, if you know the physical and chemical properties of one element in a group, you can predict the physical and chemical properties of any other element in the same group. The periodic table allows you to relate the reactivity tendencies of elements to their atomic structure. You can also predict which elements can form ionic or covalent bonds.

4.1 PERIODIC TABLE

One of the most important activities is the search for order. A large number of observations or objects can be arranged into groups according to common features they share, it becomes easier to describe them. After the discovery of atomic number by Moseley in 1913, it was noticed that atomic number could serve as a base for systematic arrangement of elements. Thus elements are arranged in the order of increasing atomic number. **A table showing systematic arrangement of elements is called periodic table.** It is based on the Periodic law that states if the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.

4.1.1 Periods and Groups of Elements.

The most commonly used form of the periodic table is shown in figure 4.1. Note that the elements are listed in order of increasing atomic numbers, from left to right and from top to bottom. Hydrogen (H) is in the top left corner. Helium (He), atomic number 2, is at the top right corner. Lithium (Li), atomic number 3, is at the left end of the second row.

The horizontal rows of the periodic table are called periods. There are varying number of elements in periods. How many periods you find in the periodic table? There are seven periods. The number of elements per period range from 2 in period 1 to 32 in period 6. First three periods are called short periods and the remaining periods are called long periods. The properties of elements within a period change gradually as you move from left to right in it. But when you move from one period to the next, the pattern of properties within a period repeats. This is in accordance to the periodic law.

International Union of Pure and Applied Chemistry (IUPAC) has recently renamed newly discovered elements and placed them on the periodic table.

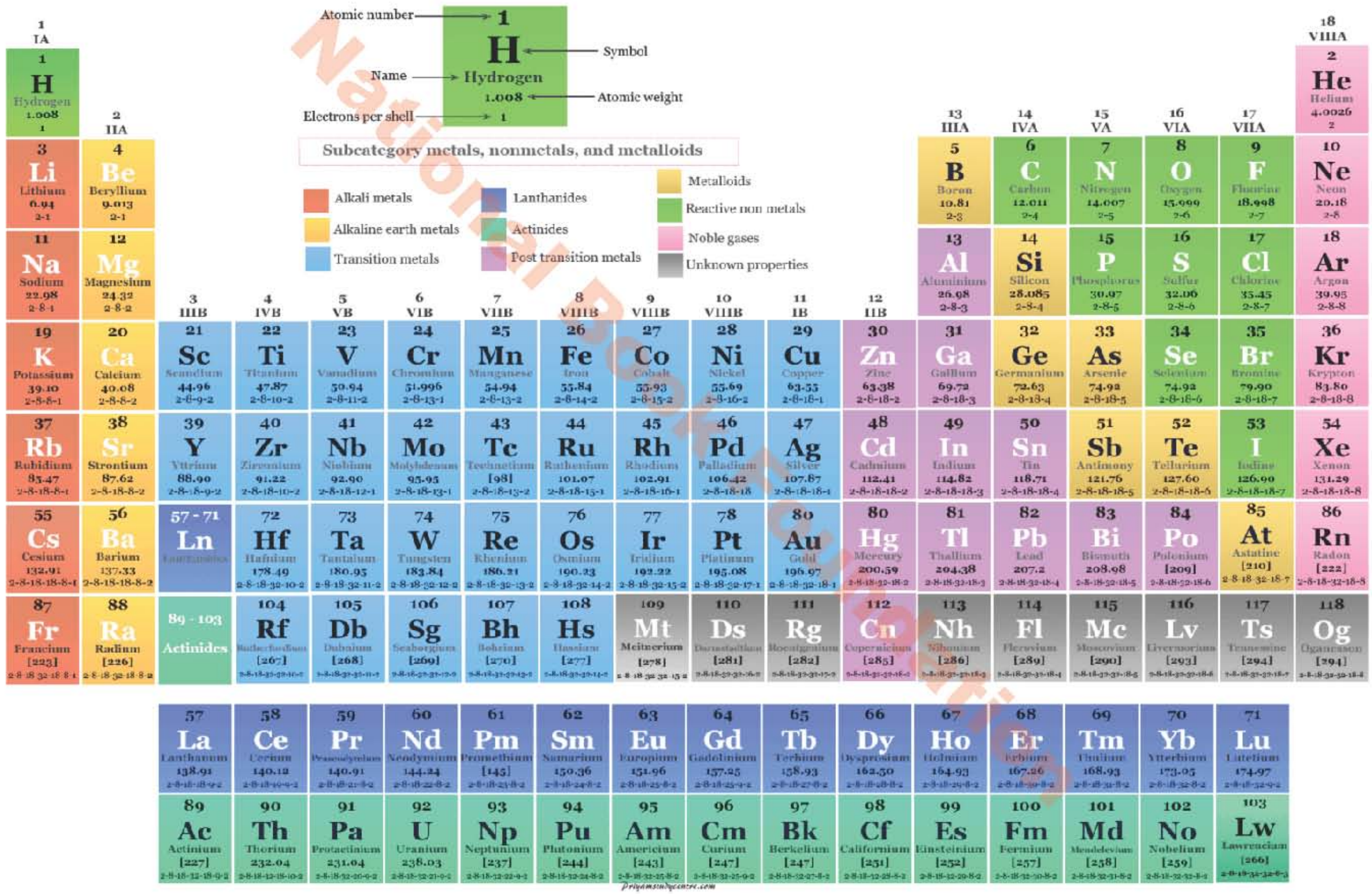
Activity 4.1

Look at the periodic table and write number of elements present in the relevant period in the table

Table Number of elements in the periods of the periodic table

Period No.	No. of elements
First	
Second	
Third	
Fourth	
Fifth	
Sixth	
Seventh	

Figure 4.1: Periodic Table of Elements



Elements that have similar properties lie in the same column in the periodic table. **Each vertical column of elements in the periodic table is called a group or family.**

Two numbering systems are often used to designate groups. You should know both. In the traditional system and the old IUPAC, the letters A and B are used. The first two groups are IA and IIA, while the last six groups are IIIA to VIIIA and the middle groups are in group B. The International Union of Pure and Applied Chemistry (IUPAC) decided that the groups would be 1-18 from left to right.

The elements in the same group have same number of valence electrons. Group number indicates the number of valence electrons in an element. For example, Group1 and Group 2 elements have 1 and 2 valence electrons respectively. In Groups 13 elements have 3, Group 14 have 4, Group15 have 5 valence electrons and so on. It is important to note that in Groups 13 to 18 (p block elements), the number of valence electrons is equal to group number minus 10.

Group A elements are called normal or representative elements. They are also called main group elements. Group B elements are called transition elements.

Names of Some Groups in the Periodic Table

Some groups of elements in the periodic table have been given group names. For example metallic elements in Group 1 are called alkali metals. Group 2 elements are called alkaline earth metals. The elements in Group 17 or VIIA are called halogens. The Group 18 or VIIIA elements are called noble gases because they do not readily undergo chemical reactions.

Recall that all elements have a unique identificcation nunber known as the atomic number or proton number. The atomic number of an element represents the number of electrons or protons present in the atom of the element. Aufbau's Principle helps in determining the order in which the electron orbitals get filled.

Electronic Configuration

According to Aufbau's principle, the order in which the orbitals fill up is as follows:

1s,2s,2p,3s,3p,4s,3d,4p,5s,4d,5p,6s,4f,5d,6p,7s,5f,6d,7p and so on.

Each orbital has a fixed capacity for the maximum number of electrons accommodated, s-orbitals have the capacity of 2 electrons, while p orbitals have the capacity for 6 electrons, d orbitals have the capacity for 10 electrons and f orbitals have the capacity for 14 electrons.

Using these concepts, we can determine the electronic configuration of the given element.

Block of an element: When you have filled all the electrons, the orbital in which the last electron is in, represents the block in which the element is placed.

Period of an element: Now, to determine the period in which the element is placed, you need to look at the principal quantum number(n) of the valence electron. This number repersents period number of element

Group of an element: To determine the group, we need to understand some rules:

(a) If the element is in s block, then the group number is equal to the number of valence electrons.

(b) If the element is in the p block, then the number of the group can be determined by the formula: (number of valence electrons + 10).

For example, the atomic number of sodium is 11.

Hence its electronic configuration is: $1s^2, 2s^2, 2p^6, 3s^1$

Since the valence electron is in the 3s subshell, sodium belongs to belongs to the s block.

The principal quantum number of the valence electron of Na is 3. Hence, it belongs to the 3rd period.

Since Na belongs to the s block, its group number is equal to a number of electrons in valence subshell s. This is equal to 1

Hence, sodium belongs to the Group 1.

Note:

we can start filling the orbitals in the order mentioned by the Aufbau principle.

1s,2s,2p,3s,3p,4s

Example 4.1: Identifying the group and period of an element

Identify the group, period, and block of following elements on the basis of electronic configuration.

1. Al (atomic number= 13)
2. K(atomic number = 19)

Problem Solving Strategy:

Write electronic configuration of element. Identify its valence shell. Remember that n value of the valence shell indicates period. Total number of electrons in the valence shells represents group number if element belongs to s block. If it belongs to p block, then group number is equal to the total number of valence + 10.

Solution:

1. Electromnic configuration of Al (atomic no. 13) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

   Valence sub- shells is 3p, so Al belongs to p block

   As n = 3, Al is present in the 3rd period.

   Total number of electrons in the valence shell = 2+1= 3

   Group number of Al= total number of electrons in the valence sub-shells + 10

   = 3+ 10

   = 13

   Hence Al belongs to Group 13

2. Electronic configuration of K (atomic no. 19) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

   Valence shells is 4s. hence K belongs to s block

   As n = 4, K is present in the 4th period.

   Total number of electrons in the valence shell = 1

   Group number of K= total number of electrons in the valence sub-shells

   = 1

   Hence K belongs to Group 1

CONCEPT ASSESSMENT EXERCISE 4.2

Identify the group and period of the following elements on the basis of electronic configurations.

(a) $^{28}_{14}Si$, (b) $^{32}_{16}S$, (c) $^{19}_{9}F$, (d) $^{40}_{18}Ar$

Example 4.2: Classifying or dividing elements into groups and periods

Electronic configuration of atoms of some elements are given below. Classify them in groups and periods.

A. $1s^2 2s^2$
B. $1s^2 2s^2 2p^3$
C. $1s^2 2s^2 2p^5$
D. $1s^2 2s^2 2p^6 3s^2$
E. $1s^2 2s^2 2p^6 3s^2 3p^5$
F. $1s^2 2s^2 2p^6 3s^2 3p^3$

Problem solving Strategy:

Remember that:

1. The elements whose atoms have similar valence shell electronic configuration belong to the same group.
2. The n value of the valence shell indicates period.
3. The elements whose atoms have same value of n for the valence shell lie in the same period.

Solution:

	II A	V A	VII A	
2	A $2s^2$	B $2s^2 2p^3$	C $2s^2 2p^5$	Period
3	D $3s^2$	F $3s^2 3p^3$	E $3s^2 3p^5$	Period

CONCEPT ASSESSMENT EXERCISE 3.1

Electronic configuration of atoms of some elements are given below. Place them into groups and periods.

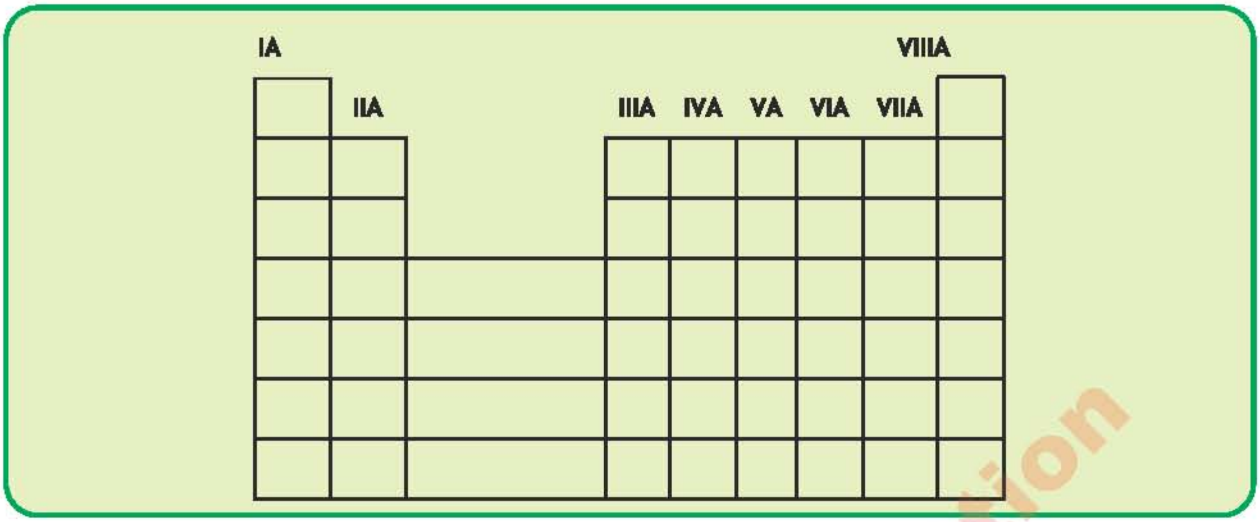
$P = 1s^2, 2s^2, 2p^2$ $\quad Q = 1s^2, 2s^2, 3p^1$

$R = 1s^2$ $\quad S = 1s^2, 2s^2$

$T = 1s^2, 2s^1$ $\quad W = 1s^2, 2s^2, 2p^6$

$X = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$ $\quad Y = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

$Z = 1s^2, 2s^2, 2p^1$



4.1.2 s and p Blocks in the Periodic Table

Group 1 and Group 2 elements contain their valence electrons in the s sub-shell. Therefore, these elements are called s-block elements. Elements in groups 13 to 18 (except He) are known as p-block elements because their valence electrons are located in the p sub-shell. Lanthanides and actinides are known as f-block elements since their valence electrons lie in f sub-shell. Figure 4.2 shows the blocks of the periodic table.

Li = 1^{s2},$2s^1$, as valence electron is in s sub-shell, Li belongs to s-block.

C=$1s^2$,$2s^2$,$2p^2$, as valence electron is in sub-shell p, C belongs to p- block.

4.2 GROUP NUMBER AND CHARGE ON AN ION

The group number of an element in the periodic table can provide information about the charge of an ion formed by an element. Valence electrons are involved in the formation of ions. The relationship between group number and ions formed by elements is based on the number of valence electrons in the element.

The group number of an s-block element in the periodic table corresponds to its number of valence electrons.

Whereas in the case of p-block elements, the number of valence electrons is equal to Group number minus 10.

Some elements tend to lose electrons. Why? Elements tend to achieve a stable electron configuration such as the noble gases. Remember that the 2 or 8 electron configuration is the most stable configuration. Elements with 1-3 electrons in their valence shell tend to lose those electrons and form +1, +2, +3 ions respectively. Elements with 5-7 electrons in their valence shell tend to gain 3, 2, 1 electrons respectively and form negatively charged ions with -3, -2, -1 charges respectively. Elements with 4 valence electrons can lose 4 electrons to form +4 ions. They can also gain 4 electrons and form -4 ions.

Group 1 (alkali metals): Group 1 elements such as lithium (Li), sodium (Na), and potassium (K) have one valence electron and belong to s block. S block elements lose electrons equal to their group number.They tend to lose this electron to form a +1 ion, also known as a mono-valent cation. For example: Lithium (Li) loses one valence electron to form Li^+ Sodium (Na) loses one valence electron to form Na^+. Potassium (K) loses one valence electron to form K^+. These elements after losing an electron acquire 8 electron configuration of a noble gas.

$Na = 1s^2, 2s^2, 2p^6, 3s^1$

$Na^+ = 1s^2, 2s^2, 2p^6$

Group 2 (alkaline earth metals): Group 2 **elements such as beryllium** (Be), magnesium (Mg), and calcium **(Ca)** have two valence electrons and are s block element. They tend to lose these two electrons to form + 2 ions, also **called** divalent cations. For example: Beryllium (Be) loses two valence electrons to form **Be^{2+}**. Magnesium (Mg) loses two valence electrons to form **Mg^{2+}**. Calcium (Ca) loses two valence electrons to form **Ca^{2+}**.

$Mg = 1s^2, 2s^2, 2p^6, 3s^2$

How many electrons Mg can lose to achieve stable electron configuration?

Magnesium will lose 2 electrons to achieve stable configuration and this no. is same as it group number i.e., 2

$Mg^{2+} = 1s^2, 2s^2, 2p^6$

Some elements tend to gain electrons to achieve noble gas configuration. For example, Group 17 (Halogens): Group **17 elements such as** fluorine (F), chlorine (Cl), and bromine **(Br)** have seven valence electrons. They tend to gain one electron to **reach** a stable octet and form **- 1 ion,** also **called a** monovalent **anion.** For example: Fluorine (F) gains one electron to form F^-. Chlorine (Cl) gains one electron to form Cl^-. Bromine (Br) gains one electron to form Br^-.

$F = 1s^2, 2s^2, 2p^5$

$F^{1-} = 1s^2, 2s^2, 2p^6$

Similarly, $Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

$Cl^{1-} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Group 16 (chalcogens): Group 16 **elements such as** oxygen (O), sulfur (S), and selenium **(Se)** have six valence electrons. They tend to gain two electrons **to reach** a stable octet and form **a -2 ion,** also **called a** divalent **anion.** For example: Oxygen (O) gains two electrons to form O^{2-}. Sulfur (S) gains two electrons to form S^{2-}.

Group 18 (precious gases): Group 18 **elements such as** helium (He), neon (Ne), and argon **(Ar)** have full valence electron shells (except helium, which has only two valence electrons). They are chemically stable and do not form ions under normal conditions. Noble gases are known for their low reactivity due to their stable electronic configuration .

Example 4.3: Obtaining the position of element in the periodic table from the electronic configuration

Find out the position of the following elements in the periodic table from the electronic configuration:

Nitrogen (atomic number: 7) (b) Oxygen (atomic number:8)

Problem Solving Strategy:

Write electronic configuration of the element. Identify the valence shell configuration, co-efficient of s or p sub-shell represents period number and total number of electrons in valence shell is equal to the group number.

Solution:

a) Electronic configuration of N = $1s^2,2s^2,2p^3$

Valence shell has configuration = $2s^2,2p^3$

Period number = 2

Number of valence electrons = 2 + 3 = 5

N belongs to p-block

So, Group number = 5 + 10 = 15

$\therefore$ Nitrogen is present in the 2^{nd} period of Group 15

b) Electronic configuration of oxygen = $1s^2,2s^2,2p^4$

Valence shell has configuration = $2s^2,2p^4$

So, Period number = 2

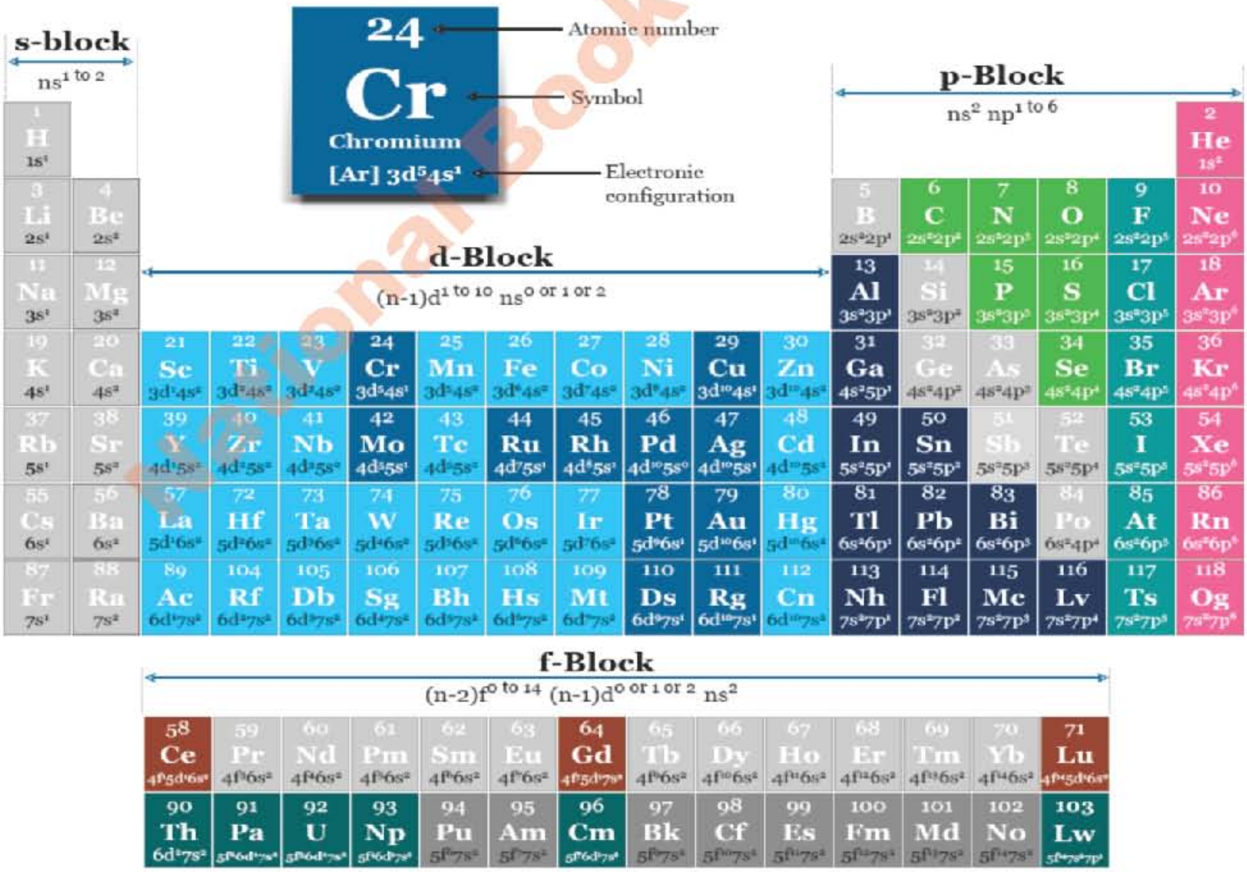
Number of valence electrons = 2 + 4 = 6

O belongs to p-block

So, Group number = 6+10 = 16

Oxygen is present in the 2^{nd} period of Group 16

Figure 4.2: Periodic Table of Elements



CONCEPT ASSESSMENT EXERCISE 4.4

Obtain the valence shell configuration of Al and S from their position in the periodic table

4.3 PERIODICITY OF PROPERTIES

There is a periodic fluctuation in the electronic configuration of the elements as the atomic number increases. Therefore, the physical and chemical properties of the elements very in a periodic manner. Elements with a similar valence shell electronic configuration are placed in the same group, one below the other. Chemical properties depend on the electronic configuration of the valence shell. Because all the elements in a given group have similar valence shell electronic configurations, they have similar chemical properties. Physical properties depend on the size of atoms. Because the sizes of atoms change gradually from top to bottom in a group. Therefore, the elements show a gradation of physical properties within the same group. In the period of the periodic table, the number of electrons in the valence shell increases gradually from left to right. Their chemical and physical properties also differ in the same way. In this section, you will learn about the variation of physical properties of certain elements within a group and across a period.

4.3.1 Shielding Effect

Figure 4.3 shows electronic configuration of Li, Be and Mg.

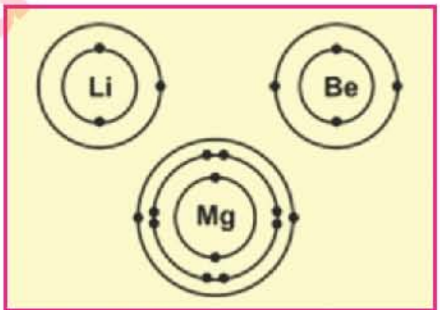


Fig 4.3 Electronic structure of Li, Be and Mg

Which atom has more shells, Be or Mg? Which atom has more electrons between the nucleus and the valence electrons, Be or Mg?

Electrons present in the inner shells cut off attractive force between the nucleus and the valence electrons.

The reduction in force of attraction between nucleus and the valence electrons by the electrons present in the inner sub-shells is called shielding effect.

Which atom has greater shielding effect, Be or Mg?

As you move from top to bottom in a group the number of electronic shells increase. So the number of electrons in the inner shell also increase. As a result shielding effect increases.

Which atom, Li or Be has greater number of shells? Which atom, Li or Be has greater number of electrons between nucleus and valence electrons?

As you move from left to right in a period the number of electrons in the inner shells remains constant, therefore, shielding effect remains constant.

Example 4.4: Identifying the element whose atoms have greater shielding effect, using periodic table

Choose the elements whose atoms you expect to have greater shielding effect.

(a) Be or Mg (b) C or Si

Problem Solving Strategy:

Look at the periodic table and find the relative position of given elements in the periodic table. Apply the trend of increasing shielding effect in a group.

Solution:

(a) Mg atoms will have greater shielding effect.

(b) Si atoms will have greater shielding effect.

CONCEPT ASSESSMENT EXERCISE 4.5

Choose the element whose atoms you expect to have smaller shielding effect.

(a) F or Cl (b) Li or Na (c) B or Al

All the physical and chemical properties of elements depend on the electronic configuration of their atoms. We now consider some properties of atoms that are affected by electronic configuration: atomic size, ionization energy, electron affinity and electronegativity. They usually increase and decrease repeatedly throughout the periodic table. That is, they show consistent changes or trends within a group or a period. These tendencies are correlated with behavior.

4.3.2 Atomic Size

The size of an atom depends on its electronic configuration. **Atomic size is the average distance between the atomic nucleus and the electronic outer shell.** Figure 4.4 shows the atomic radii of the main group elements. Figure 4.4 shows the variation of atomic radii within a period and within a group. You can see two general trends in atomic radii.

1) The atomic radius decreases in each period as you move across the period. This is because as you move from one element in the sequence to the next, to the right of it. Another electron is added to the same valence shell. At the same time, the positive charge of the core also increases by one. The attraction of the nucleus to the electron in the valence shell increases. Therefore, the size of the shell

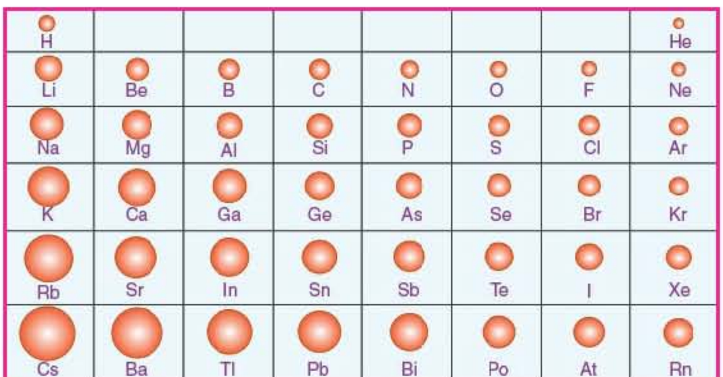


Figure 4.4: Atomic sizes of the main group elements

and the radius of the atom decreases. For example, going from lithium to beryllium, the atomic size decreases. You can understand this from the electronic configuration of the valence shell of Li ($2s^1$) and B ($2s^2$). Moving from Li to Be, the number of shells does not change, but the atomic number increases from 3 to 4. Therefore, the strength of the nucleus on the valence shell electron increases. Therefore, the atomic radius decreases.

2) Atomic radius increases in each main group as you move down the element group. This is because the size of an atom is determined by the size of its valence shell. As you move down the group to the next lower element, the atom has an additional shell of electrons. This increases the radius of the atom. For example, going from Li to Na, the atomic radius increases. Consider the electron configuration of Li ($1s^2\ 2s^1$) and Na ($1s^2$, $2s^2$, $2p^6$, $3s^1$). A new electron shell is added, increasing the size of the atom.

Example 4.5: Identifying the element that has greater atomic radius

Choose the element whose atom you expect to have larger atomic radius in each of the following pairs.

(a) Mg, Al (b) C, Si

Problem Solving Strategy:

Remember that the larger atom in any:

(a) Period lies further to the left in the periodic table.

(b) Group lies closer to the bottom in the periodic table.

(c) Check the periodic table and choose the element.

Solution:

(a) The larger atom is Mg

(b) The larger atom is Si

CONCEPT ASSESSMENT EXERCISE 4.6

Using the periodic table but without looking at the figure 4.4, choose the element whose atom you expect to have smaller atomic radius in each of the following pairs.

(a) O or S (b) O or F

4.3.3 Ionization Energy

Ionization energy is an important property of atoms that explains cation formation. **"Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom".**

$$M(g) + \text{ionization energy} \longrightarrow M^+(g) + \acute{e}$$

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H 1312							He 2372
2	Li 520	Be 899	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
3	Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1521
4	K 419	Ca 590	Ga 579	Ge 762	As 947	Se 941	Br 1140	Kr 1351
5	Rb 403	Sr 549	In 558	Sn 709	Sb 834	Te 869	I 1008	Xe 1170
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At 926	Rn 1037

Figure 4.5 Ionization energies of the main group elements

Ionization energy is a measure of the extent to which the nucleus attracts the outermost electron. A high value of ionization energy means stronger attraction between the nucleus and the outermost electron. Whereas a low ionization energy indicates a weaker force of attraction between the nucleus and the outermost electron. Figure 4.5 shows the ionization energies of the main group elements. Values are given in units of kJ/mole^{-1}or kJ/mole.

Trends in ionization energy values.

The value of the ionization energy decreases from top to bottom in the group. This is because the shielding effect of the atoms increases down the group. Greater shielding effects result in a weaker attraction of the valence electrons to the nucleus. So they are easier to remove. This leads to a decrease in ionization energy from top to bottom in the group. Which atom has a greater shielding effect, Li or Na? As you move from left to right in the period, the shielding effect remains unchanged. But little by little the nuclear charge increases. The stronger attraction between the nucleus and the valence electron increases. As a result, the ionization energy increases from left to right in a period. Which atom has the higher ionization energy, Li or Be?

Example 4.6: Identifying the element that has smaller ionization energy

Choose the element whose atom you expect to have smaller ionization energy in each of the following pairs.

(a) B,C (b) N, P

Problem Solving Strategy

Remember that ionization energy:

(a) Increases across a period. The element that has smaller ionization energy will be further to the left in the periodic table.

(b) Decreases from top to bottom in a group. The element that has smaller ionization energy will correspond to the element closer to the bottom.

(c) Check the periodic table to choose the element.

Solution:

(a) The atom with the smaller ionization energy is B

(b) The atom with the smaller ionization energy is P.

CONCEPT ASSESSMENT EXERCISE 4.7

Which atom has the smaller ionization energy?

(a) B or N (b) Be or Mg (c) C or Si

4.3.4 Electron Affinity

Electron affinity explains the anion formation. **Electron affinity is defined as the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uni negative gaseous ion.**

$$X(g) + e^{-} \longrightarrow X^{-}(g) + \text{electron affinity}$$

Figure 4.6 shows electron affinities of main group elements.

Factors affecting electron affinity are nuclear charge, atomic radius and shielding effect.

As you move from left to right through a period, electron affinity generally increases. This is due to an increase in nuclear charge and a decrease in atomic radius, which binds the extra electron more tightly to the nucleus. But the shielding effect remains constant in each cycle. Therefore, the alkali metals have the lowest and the halogens the highest electron affinities in each period.

Electron affinity decreases from top to bottom in a group. This is due to an increase in the shielding effect. Due to the increased shielding effect and increase in atomic radius, the added electron binds less tightly to the nucleus. As a result, less energy is released.

H -73								He 0
Li -60	Be 0		B -27	C -122	N +7	O -141	F -328	Ne 0
Na -53	Mg 0		Al -44	Si -134	P -71.7	S -200	Cl -349	Ar 0
K -48	Ca 0		Ga -29	Ge -120	As -77	Se -195	Br -325	Kr 0
Rb -47	Sr 0		In -29	Sn -121	Sb -101	Te -190	I -295	Xe 0
Cs -45	Ba 0		Tl -30	Pb -110	Bi -110	Po -180	At -270	Rn 0

Figure 4.6 electron affinities of main group elements

There are several exceptions to the general trend of election affinity values. You will learn reasons for it in higher grade.

4.3.5 Electronegativity

Electronegativity is the ability of an atom to attract electrons toward itself in a chemical bond. Figure 4.7 shows as a scale of electronegativities of the elements devised by Linus Pauling. The American chemist Linus Pauling devised a method for calculating the relative electronegativities of elements.

H 2.1							He
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne 2.1
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar 3.0
K 0.5	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 2.1
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9						

Figure 4.7 the electronegativities of elements.

Activity 4.2

Determining the general trends in the electronegativities

You will need:

- Figure 4.7

Carry out the following:

1. Move across the second period from left to right and note down the variation in electronegativity values.
2. Move across the 3rd period from left to right and note down the variation in electronegativity values.
3. Make generalization about the variation in electronegativites across a period and write reason.
4. Move from top to bottom in Groups IA and IIA and note down the variation in electronegativites value.

5. Move from top to bottom in Groups VIA and VIIA and note down the variation in electronegativities value.

Make generalization about the trend in electronegativity values in a group. Give reason.

4.4 CHARACTERISTIC PROPERTIES

Characteristic properties of an element in a given group are based on periodicity and chemical reactivity. For example, in Group 1 (alkali metals) such as lithium, sodium, potassium, are highly reactive metals. They have general electron configuration ns^1. Their reactivity trend increases as you move down the group. Lithium, being at the top of the group is the least reactive metal among alkali metals. As you move down the group, the atomic size increases and the outer most electron is further from the nucleus, leading to be lost easily. This leads to increased reactivity. So, sodium is more reactive than lithium. Which is more reactive sodium or potassium ?

Similar trend is observed in Group 2 (alkaline earth metals). Which is more reactive Mg or Ca ?

4.4.1 Metallic Character

Metallic nature refers to a property of elements in the periodic table that determines how easily they can lose electrons and form positive ions (cations). Elements with metallic character have a strong tendency to lose electrons and easily form cations. The metallic character of an element is affected by its position in the periodic table.

Metallic character increases as you move down a group in the periodic table. This is primarily due to addition of new electronic shells. The outermost electrons are farther from the nucleus and experiences weaker attractive forces, making it easier for them to be lost. This promotes metallic character.

Metallic character decreases as you move across a period from left to right in the periodic table. This is because effective nuclear charge increases across a period, while the number of shells remains the same. The stronger attractive forces make it more difficult for valence electrons to be lost.

Example 4.6: Identifying the element that has higher metallic character.

Choose the element you expect to have higher metallic character in each of the following pairs.

(a) Na or K

(b) Na or Mg

Problem Solving Strategy

Remember that metallic character :

(a) Increases down the group. The element that has higher metallic character will be closer to the bottom.

(b) Decreases across a period. The element that has higher metallic character is further to the left.

(c) Check the periodic table to choose the element.

Solution:

(a) K

(b) Mg

CONCEPT ASSESSMENT EXERCISE 4.8

Which element has lower metallic character?

(a) Li or K

(b) Mg or Ca

(c) Compare and contrast ionization energy and electron affinity

4.4.2 Reactivity

The capability of an element to react with other elements to form new compounds is called it reactivity. Reactivity of elements generally increases as you move down a group. This is due to the increase in atomic size. The outermost electrons are farther from the nucleus and experience weaker attractive forces, making it easier for them to participate in chemical reaction.

Reactivity tends to very across a period. Elements on the left side of a period (Group 1 and 2) are highly reactive due to their strong tendency to lose electrons and form positive ions. Elements on the right side of a period (Group 16 and 17) are highly reactive as well but tend to gain electrons to form negative ions.

4.4.3 Density

Density of elements generally increases as you move down a group. This is due to the increasing atomic mass and the larger size of atoms. As the number of protons and neutrons in the nucleus increases, the atomic mass increases. This results in higher density.

Density can wary across the period. In general, density tends to increase from left to right until it reaches a maximum around the middle of the period, and then it starts to decrease.

4.4.4 Characteristic Properties of Alkali Metals

Some characteristic properties of alkali metals are as follows:

(a) Highly reactive metals:

Alkali metals are highly reactive metals in the periodic table. They readily lose valence electron to form a +1 cations. This trend increases down the group.

(b) Softness and low density:

Alkali metals have low densities and are relatively soft, which allows them to be easily cut with a knife. This trend increases down the group the group.

(c) They are excellent conductors of electricity and heat.

(d) They have low melting points.

(e) They are highly reactive and monovalent elements.

(f) They react with H_2O to give H_2 and alkali metal hydroxides.

Which is more soft Na or K?

4.4.5 Prediction of properties of other elements in Group 1

In Group 1 lithium, sodium and potassium are a collection of relatively soft metals showing a trend in melting point and reaction with water.

The metals in group I are called alkali metals.

- They are very soft.
- Their melting and boiling points decrease down the group.
- When alkali metals react with water, they produce a metal salt and hydrogen.

  metal + water ——> salt + hydrogen
- The alkali metals become more reactive down the group.

Activity 4.5

Predict the properties of other elements in Group I, from the data given above.

The element after Potassium is Rubidium and you can predict that its reaction with water will be much more violent. We can also predict that Rubidium will have a lower melting and boiling point than the three elements above it. And the elements below Rubidium will be even more reactive and have very low melting and boiling points. It will also react with water to for salt and hydrogen.

4.4.6 Position of Unknown Element in the Periodic Table

You can place an unknown element accurately at appropriate position in the periodic table, and can predict about its properties.

The electronic configuration of an element strongly influences its chemical behaviour. Elements within the group have similar electronic configuration, and therefore similar properties. By examining the electronic configuration of unknown element and comparing it to the known elements in the periodic table, it`s likely position in the periodic table can be determined.

The periodic table allows for the identification of trends and patterns across periods and groups. These trends include variations in atomic size, ionization energy, electron affinity, electronegativity, reactivity, and other properties. By analyzing these trends, it becomes possible to estimate the general properties of unknown element and make guess about its position in the periodic table.

Example 4.7: Identifying position of an unknown element in the periodic table

Suppose you have an unknown element having atomic number 19, and you want to determine its position in the periodic table.

Problem Solving Strategy

1. Write its electronic configuration.
2. Use valence electronic configuration to locate its position i.e., find its group and period.

Solution:

Electronic configuration : $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

Valence shell electronic configuration is $4s^1$, which shows it is an alkali metal, because, Group 1 elements have one electron in valence sub- shell s. As n value of valence sub-shell is 4, this unknown element must lie in the 4^{th} period in the periodic table.

From its position in the periodic table, you can predict its properties. For example will it possess higher or lower melting point, density, reactivity, etc. than the element above or below it.

4.5 TRANSITION ELEMENTS

Elements located in d-block (Group 3 to 12) in the periodic table are called transition elements. These elements exhibit several characteristic properties, which set them apart from other elements. Some of their properties are as follows.

1. High Density

   Transition elements generally possess high densities due to their higher atomic masses and closely packed structures. For example, iron(Fe) has a density of 7.87 g/cm^3, tunguston(W) has a density of 19.3g/cm^3.

2. High Melting Points

   Transition elements have high melting points. This is because their metallic bonding is stronger, which in tern is due to the presence of partially filled d-sub shells. For example tunguston has a melting point of 3422 ^oC, platinum(Pt) has a melting point of 1768 ^oC.

3. Variable Oxidation States

   Transition elements exhibit multiple oxidation states. This is because of d-sub shell can also participate in bonding along with s-sub shell. For example, iron(Fe) has oxidation states +2 and +3, copper(Cu) has oxidation states +1 and +2.

4. Coloured Compounds

   Transition elements often exhibit vibrant colours. For example, copper compounds appear blue or green, chromium compounds are often red or green.

5. Catalysts for Industrial Processes

   Transition metals and their compounds are widely used as catalyst in various industrial processes. For example,

   (a) Iron is used in the Haber Process for the synthesis of ammonia.

   (b) Platinum and palladium are used as catalyst in catalytic converters to reduce harmful emissions in automobiles and industrial units.

   (c) Nickel is used as catalyst in the manufacture of margarine.

   (d) Platinum is used as catalyst in the contact process for the manufacture of sulphuric acid.

4.6 LANTHANIDES & ACTINIDES

Lanthanides also known as "rare earth elements" are series of 14 elements located at the bottom of the periodic table. They include elements with atomic number 57 to 71.

Actinides are another series of 14 elements located just below lanthanides. They include elements with atomic number 89 to 103.

4.7 HALOGENS

The elements in group 17 (or Gruop VII-A) are called halogens. The name halogen is derived from the Greek words "halous" meaning salt and "gen" meaning former. Halogens include fluorine, chlorine, bromine, iodine, astatine, and tenessine (astatine and tenessine are radio- active elements. Little is known about their properties). All halogens are reactive non-metals and exist as diatomic molecules.

4.7.1 Appearance of halogens

They all exist as diatomic coloured molecular substances. The colour of halogen become darker as you go down the group. At room temperature and pressure(RTP) fluorine(F_2) exist as pale yellow gas, chlorine(Cl_2) as yellow-green gas, bromine(Br_2) as red-brown liquid and iodine(I_2) as grey-black solid. Iodine easily turn into a dark purple vapours on warming.

Electronic Configuration

Halogens possess 7 electrons in their valence shell. They have general electronic configuration ns^2np^5. They need only one electron to complete their valence shell. Consequently, they tend to gain one electron to form univalent negative ions, F^-, Cl^-, Br^-, I^-.

Density of halogens

As you move down the group the number of electrons and protons increases, the size of the atom increases and the volume increases. However, the increase in mass exceeds the increase in volume, so the density, which is mass per unit volume, increases in general. Also fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid . So, the forces of attraction between molecules increase down the group. Solid iodine has molecules that are highly attracted and tightly packed together than bromine. Therefore, as you go down the group of halogens, the forces of attraction increase and the density of the halogen increases.

Densities of halogens

Halogen	Density (g/cm^3 at 25 °C)
Fluorine	0.0017
Chlorine	0.0032
Bromine	3.1028
Iodine	4.933

Reactivity of halogens

The reactivity of halogens is directly related to their ability to gain an electron and form a halide ion (fluoride ion F^-, chloride ion Cl^-, bromide ion Br^-, iodide ion I^-) when they react with other elements. Fluorine has the greatest tendency to gain electrons and form a halide ion, making it the most reactive halogen. As you move down the group, the electronegativity of the halogens decreases. This leads to a decrease in reactivity. Which halogen is the least reactive? Bromine or iodine. Because halogens have a strong tendency to gain electrons, they have a strong oxidizing power, and this power decreases down the group. Thus, the order of decreasing oxidizing power is.

$$F_2 > Cl_2 > Br_2 > I_2$$

Displacement reactions of halogens

Oxidizing power of F_2 is the highest and that of I_2 is lowest. Due to the relative strength as oxidizing agent, it is possible for a free halogen to oxidize or displace the ion of halogen next to it in the group from their aqueous solutions. This means F_2 can oxidize and displace all the halide ions to free halogen. For example,

$$F_{2(g)} + 2KCl_{(aq)} \rightarrow 2KF_{(aq)} + Cl_{2(g)}$$
$$F_{2(g)} + 2KBr_{(aq)} \rightarrow 2KF_{(aq)} + Br_{2(g)}$$
$$F_{2(g)} + 2NaI_{(aq)} \rightarrow 2NaF_{(aq)} + I_{2(g)}$$

Similarly Cl_2 can oxidize Br^- and I^- ions. But I_2 can not oxidize any halide ion.

Hydrogen halides and their thermal stabilities

Halogens react with hydrogen to form hydrogen halides.

$H_2 + X_2 \rightarrow HX$ Where X =F_2, Cl_2, Br_2, I_2

The strength of the hydrogen-halogen bond is related to the electronegativity difference between the hydrogen and halogen atoms. A larger electronegativity difference results in a stronger bond. As we move from HF to HI, the electronegativity difference between the hydrogen and halogen atoms decreases, resulting in weaker bonds in HCl, HBr, and HI. So, the relative thermal stability of hydrogen halides gradually decreases from HF to HI.

Consequently, the energy needed to break H-X decreases in the following orders

HF> HCl> HBr > HI

Prediction of properties of elements in Group VIIA or Group 17

The elements present in Group 17 or VIIA are called halogens. They are poisonous non-metals that have low melting and boiling points that increase down the group. As a result of this increasing boiling and melting points, the state of the halogens at room temperature, changes from gas to liquid to solid down the group (fluorine and chlorine, the 1st and 2nd halogens, are a gas; bromine, the 3rd halogen is a liquid; and iodine, the 4th halogen, is a solid). The colours of halogens also get darker from top to bottom.

Activity 4.6

Predict the properties of other elements in Group VII, from the given data given above.

From this data you can predict how the halogens will behave up and down the group. Astatine, the fifth halogen, will have high melting and boiling points so will be solid at room temperature, and will have a very dark colour.

CONCEPT ASSESSMENT EXERCISE 4.9

Which of the following displacement reactions will occur?

1. $Cl_{2(g)} + 2NaF_{(aq)} \longrightarrow 2NaCl_{(aq)} + F_{2(g)}$
2. $Br_{2(g)} + 2KI_{(aq)} \longrightarrow 2KBr_{(aq)} + I_{2(g)}$
3. $I_{2(g)} + 2KBr_{(aq)} \longrightarrow 2KI_{(aq)} + Br_{2(l)}$
4. $Cl_{2(g)} + 2KBr_{(aq)} \longrightarrow 2KCl_{(aq)} + BR_{2(l)}$
5. $Cl_{2(g)} + 2NaI_{(aq)} \longrightarrow 2NaCl_{(aq)} + I_{2(s)}$

4.8 NOBLE GASES

Noble gases, also known as inert gases, are a group of chemical elements found in Group18 (or Group VIII-A) of the periodic table. They have general electron configuration ns^2,np^6 except He, which has $1s^2$. They are characterized by unique properties. They are odorless, colorless mono-atomic gases and possess very low reactivity with other elements. This low reactivity is due to the presence of a complete valence shell, which makes them stable and unlikely to form chemical bonds with other elements under normal conditions. Noble gases include elements: Helium (He) neon (Ne) argon (Ar) crypto (kr) xenon (Xe) radon (Rn), and oganesson (Og).

DO you know?

Due to their non-reactive nature, noble gases are used in many ways, such as in lighting (e.g. neon signs), refrigeration systems and welding. They are also used in special applications, including filling gas exhaust lines and as a shielding gas in certain industrial processes.

Table 4.1: Electronic Configuration of Nobal Gasses

Element	Atomic Number	Electronic Configuration
Helium	2	$1s^2$
Neon	10	$1s^2 2s^2 2p^6$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
Krypton	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
Xenon	54	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$ $4p^6 4d^{10} 5s^2 5p^6$

4.9 COMPARISON OF GENERAL PHYSICAL PROPERTIES OF METALS AND NON-METALS:

Thermal Conductivity:

Metals generally have high thermal conductivity, which means they can conduct heat easily. On the other hand, non-metals tend to have poor conductivity, making them less efficient at conducting heat.

Electrical Conductivity:

Metals are good conductor of electricity, because they have free electrons that can move freely in the metal lattice. Non-metals, with few exception such as graphite, are poor conductor of electricity because they lack free electrons.

Adaptability:

The metals are malleable and ductile. So, they can be hammered, drawn into wires or transformed into thin sheets without breaking. This property is due to metallic bonds which allow atoms exchange easily under pressure. Non-metals are neither malleable nor ductile rather they are brittle.

Melting Points and Boiling Points

Metals generally have high melting points and boiling points due to strong metallic bonds that require a lot of energy to break. On the other hand non-metals often have lower melting points and boiling points because their atoms and molecules are held by weaker bonds such as covalent bonds, van der Waals bonds, or hydrogen bonds that require less energy to break.

CONCEPT ASSESSMENT EXERCISE 4.10

Compare the general properties of metals and non-metals

KEY POINTS

- When elements are arranged in the order of their increasing atomic number, their properties are repeated in a periodic manner.
- A horizontal row of elements in the periodic table is called a period.
- A column of elements in the periodic table is called a group or a family.
- Group IA and IIA elements are called s-block elements, since s sub-shell fills in these elements.
- Elements in the same group possess similar chemical properties.
- Elements in group IIIA to VIIIA are called p-block elements, because filling of valence p sub-shell occurs in these elements.
- The length of a period in the periodic table depends on the type of sub-shell that fills.
- The decrease in force of attraction between nucleus and the valence electron by the electrons present in the inner sub-shells is called shielding effect.
- The size of atom is the average distance between the nucleus of an atom and the outer electronic shell.
- The atomic radii decrease from left to right in a period. Whereas these increase from top to bottom in a group.
- Ionization energy is the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom.
- Electron affinity is the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uni-negative gaseous ion.

References for additional information

- B.Earl and LDR Wilford, Introducion to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).
- Lawarie Ryan, Chemistry for you.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Number of periods in the periodic table are:

(a) 8 (b) 7

(c) 16 (d) 5

(ii) Which of the following groups contain alkaline earth metals?

(a) 1A (b) IIA

(c) VIIA (d) VIIIA

(iii) Which of the following elements belongs to VIIIA?

(a) Na (b) Mg

(c) Br (d) Xe

(iv) Main group elements are arranged in ________ groups.

(a) 6 (b) 7

(c) 8 (d) 10

(v) Period number of $^{27}_{13}Al$ is:

(a) 1 (b) 2

(c) 3 (d) 4

(vi) Valence shell electronic configuration of an element M (atomic no. 14) is:

(a) $2s^2,2p^1$ (b) $2s^2,2p^2$

(c) $2s^2,2p^3$ (d) $3s^2,3p^2$

(vii) Which of the following elements you expect to have greater shielding effect?

(a) Li (b) Na

(c) K (d) Rb

(viii) As you move from right to left across a period, which of the following does not increase:

(a) electron affinity (b) ionization energy

(c) nuclear charge (c) shielding effect

(ix) All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:

(a) high ionization energies (b) relatively greater atomic sizes

(c) similar electronic configuration (d) decreased nuclear charge

2. **Give short answer.**

   (i) Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.

   (ii) Define halogens.

   (iii) Which atom has higher shielding effect, Li or Na?

   (iv) Explain why, Na has higher ionization energy than K?

   (v) Alkali metals belong to S-block in the periodic table, why?

3. Arrange the elements in each of the following groups in order of increasing ionization energy:

   (a) Li, Na, K (b) Cl, Br, I

4. Arrange the elements in each of the following in order of decreasing shielding effect.

   (a) Li, Na, K (b) Cl, Br, I (c) Cl, Br

5. Specify which of the following elements you would expect to have the greatest electron affinity.

   S, P, Cl

6. Electronic configuration of some elements are given below, group the elements in pairs that would represent similar chemical properties.

   $A = 1s^2,2s^2$ $B = 1s^2,2s^2,2p^6$ $C = 1s^2,2s^2,2p^3$

   $D = 1s^2$ $E = 1s^2,2s^2,2p^6,3s^2,3p^3$ $F = 1s^2,2s^1$

   $G = 1s^2,2s^2,2p^6,3s^1$ $H = 1s^2,2s^2,2p^6,3s^2$

7. Arrange the elements in groups and periods in Q. No. 6.

8. For normal elements, the number of valence electrons of an element is equal to the group number. Find the group number of the following elements.

   $^{27}_{13}Al$, $^{32}_{16}S$, $^{39}_{19}K$, $^{8}_{6}O$

9. Write the valence shell electronic configuration for the following groups:
   a. Alkali metals
   b. Alkaline earth metals
   c. Halogens
   d. Noble gases

10. Write electron dot symbols for an atom of the following elements
   (a) Be (b) K (c) N (d) I

3. Write the valence shell electronic configuration of the atoms of the following elements.
   (a) An element present in period 3 of Group VA
   (b) An element present in period 2 of Group VIA

4. Copy and complete the following table:

Atomic number	Mass number	No. of protons	No. of neutrons	No. of electrons
11			12	
		14	15	
	47		25	
	27			13

13 In which block, group and period in the periodic table where would you place each of the following elements with the following electronic configurations?
   (a) $1s^2,2s^1$
   (b) $1s^2,2s^2,2p^5$
   (c) $1s^2,2s^2,2p^6,3s^2$
   (d) $1s^2$

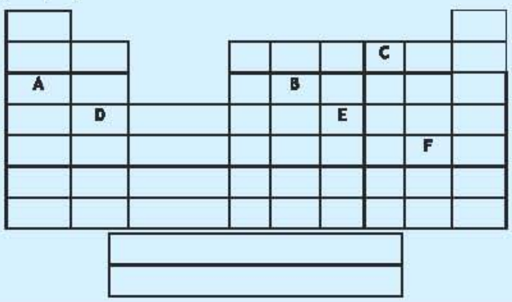
THINK TANK

14. What types of elements have the highest ionization energies and what types of elements have the lowest ionization energies. Argue.

15. i. Two atoms have electronic configuration $1s^2,2s^2,2p^6$ and $1s^2,2s^2,2p^6,3s^1$. The ionization energy of one is 2080kJ/mole and that of the other is 496kJ/mole. Match each ionization energy with one of the given electronic configuration. Give reason for your choice.

   ii. Use the second member of each group from Group IA, IIA and VIIA to judge that the number of valence electron in an atom of the element is the same as its group number.

   iii. Letter A, B, C, D, E, F indicates elements in the following figure:



a. Which elements are in the same periods?

b. Write valence shell electronic configuration of element D.

c. Which elements are metals?

d. Which element can lose two electrons?

e. In which group E is present?

f. Which of the element is halogen?

g. Which element will form dipositve cation?

h. Write electronic configuration of element E

i. Which two elements can form ionic bond?

j. Can element C form C_2 molecule? Interpret.

k. Which element can form covalent bonds?

l. Is element F a metal or non-metal?

16. Electronic configurations of four elements are given below:

(a) $1s^2,2s^1$ (b) $1s^2,2s^2,2p^5$ (c) $1s^2,2s^2,2p^6,3s^2$ (d) $1s^2$

Which of these elements is

i) An alkali metal

ii) An alkaline earth metal

iii) A noble gas

iv) A halogen

17. Argue in what region of the periodic table you will find elements with relatively

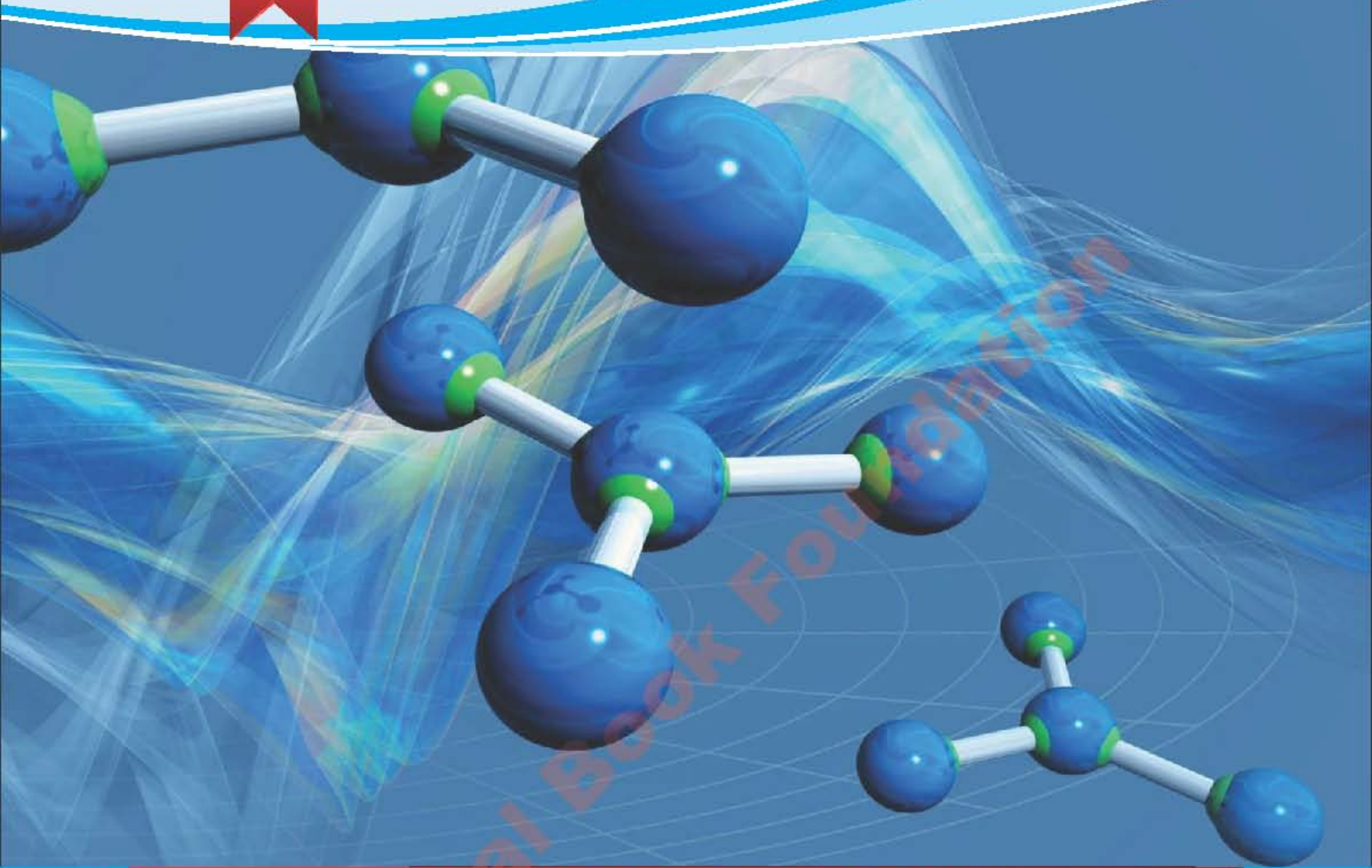
a) high ionization energies

b) low ionization energies

PROJECT

Prepare 3D model of the periodic table (Group Activity)

UNIT 05



CHEMICAL BONDING

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Describe noble gas electronic configuration, octet and duplit rules help predict chemical properties of main group elements.
- Compare between the formation of cations and anions.
- Account for the electropositive and electronegative nature of metals.
- Define ionic, covalent, coordinate covalent and metallic bonds.
- Differentiate between ionic compounds and covalent compounds. (the following points to be included in the respective definitions: a. Ionic bond as strong electrostatic attraction between oppositely charged ions. b. Covalent bond as strong electrostatic attraction between shared electrons and two nuclei. C. Metallic bonds as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations.

- Explain the properties of compounds in terms of bonding and structure.
- Compare properties and use of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.
- Interpret the strength of forces of attraction and their impact on melting and boiling points of ionic and covalent compounds.
- Justify the availability of free charged particles (electrons or ions) for conduction of electricity in ionic compounds (solid and molten) covalent compounds and metallic bonds.
- Recognize some substances can ionize when dissolved in water (e.g. acids dissolve in water and conduct electricity).
- Justify the suitability of usage of graphite, diamond and metals for industrial purposes (some example may include; a. graphite as lubricant or an electrode. b. diamond in cutting tools. c. metals for wires, and sheets).
- Draw the structure of ionic and covalent compounds along with their formation. (some examples may include: a. ionic bonds in binary compounds such as NaBr, NaF, $CaCl_2$, using dot-and-cross diagrams and Lewis-dot structures. b. simple molecules including H_2, Cl_2, O_2, N_2, H_2O, CH_4, NH_3, HCl, CH_3OH, C_2H_4, CO_2, HCN, and similar molecules using dot-and-cross diagrams and Lewis -structures.

INTRODUCTION

All the matter in this world is composed of almost entirely compounds and their mixtures. Human, animal and plant bodies, rocks, soil, petroleum, coal etc. are all complex mixtures of compounds. In compounds different kinds of atom are bounded together. Few elements also consist of unbounded atoms. For instance, helium, neon, argon, xenon and krypton present in the atmosphere consist of unbounded atoms. The manner in which various atoms are bonded together has a profound effect on the properties of substances.

Some substances are hard and tough, others are soft and flexible why ? Resins are widely used to paint dams, bridges, buildings and automobiles. What makes them sticky ? How do adhesives such as glue bind two surfaces together ? What is the nature of such linkages ? The answer lies in the nature of bonding and structure of their molecules. Therefore, to understand the behaviour of various substances, you must understand the nature of chemical bonding and structure of molecules.

5.1 WHY DO ATOMS REACT?

There are eight groups of normal elements (IA, IIA, IIIA, IVA, VA, VIA VIIA VIIIA) in the periodic table. Group VIIIA consist of the noble gases or zero group elements because they are all very stable and chemically inert under ordinary condition. They exist in atomic form in the atmosphere. They have general electronic configuration = ns^2, np^6 (8 electrons in valence shell) except He ($1s^2$). These noble gases have completely filled valence shells (s and p subshells). Their octet is complete, so they do not participate in ordinary chemical reactions and are called inert gases. They have eight electrons in their valence shell, except He, which has two electrons in its valence shell.

$_2He = 1s^2$

$_{10}Ne = 1s^2, 2s^2, 2p^6$

$_{18}Ar = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

In 1916 a chemist G. N. Lewis used the concept of octet (eight electrons) and duplet (2 electrons) electronic rule to explain the reactivity and stability of molecules.

Octet Rule

The octet rule states that an atom is most stable when its valence shell contains eight electrons. This principle is derived from the observation that atoms of the major group elements tend to participate in chemical bonding in the form of eight electrons per atom in the resulting molecule. This rule only applies to the major group element. The chemical behaviour of the main group elements can be predicted with the help of the octet rule. This is because the rule only involves s and p electrons. Molecules such as oxygen, nitrogen, and carbon follow the octet principle. Hydrogen, helium, and lithium follow the duplet rule because their electrons lie in s orbital.

$_{11}Na = 1s^2, 2s^2$, sp6, $3s^1$ (unstable, reactive, incomplete octet) Loss of one electron

$Na^+ = 1s^2, 2s^2, 2p^6$ which is same as that of Ne

$_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ (unstable, reactive, incomplete octet) $_{18}Cl^{-1} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ which is same as that of $_{18}Ar$

Duplet rule

The tendency of atoms to acquire two electronic configuration in their outermost shell during bond formation is called duplet rule. They attain electronic configuration like Helium.

For Example

$_3Li = 1s^2, 2s^1$ lose 1 electron to form $Li^+ (1s^2)$

$_4Be = 1s^2, 2s^2$ loses two electrons to form $Be^{+2} (1s^2)$

Helium has two electrons in its valence shell and is also chemically inert. Some elements that are close to He on the periodic table tend to achieve two electronic configuration in their valence shell. For example, hydrogen, lithium and beryllium etc. tend to achieve two electron configuration in the valence shell.

5.2 CHEMICAL BONDS

Atoms combine to form various types of substances. But what holds them together? Fundamentally, some forces of attraction hold atoms together in substances. These forces are called chemical bonds. Basically the forces of attraction that lead to chemical bonding between atoms are electrical in nature. Electronic structure of an atom helps us to understand how atoms are held together to form substances. Atoms other than the noble gases have a tendency to react with other elements. These elements are reactive because they tend to gain stability by loosing or gaining electrons. When atoms gain or lose electron they acquire the configuration of next noble gas element. The tendency of metal atoms to lose electrons is called electropositivity. Where as the tendency of non-metal atoms to gain electrons is called electronegativity. So, metals are electropositive and non-metals are electronegative elements.

Atoms can also acquire the configuration of next noble gas element by sharing electrons.

Electropositive and Electronegative Elements

Metals are electropositive in nature because all metal atoms lose electrons from their outermost shell in order to become stable and become positively charged. They have low ionization energy and low electronegativity allowing them to easily lose electrons. Therefore, they can form positive ions by losing electrons.

Example: $Na \rightarrow Na^{+} + e^{-}$

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

Non-metals are electronegative in nature because all non-metals gain electrons in order to become stable and hence become negatively charged. They have high electronegativity and have high electron affinity. So they can easily form negative ions by gaining electrons. For example:

$F + e^{-} \rightarrow F^{-}$

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

5.3 TYPES OF BONDS

Depending on the tendency of an atom to lose or gain or share electrons, there are two types of bonds:

1. Ionic bonds
2. Covalent bonds

5.3.1 Ionic Bonds

Ionic bonds are formed between two atoms, when one atom loses electron to form cation and the other atom gains this electron to form anion.

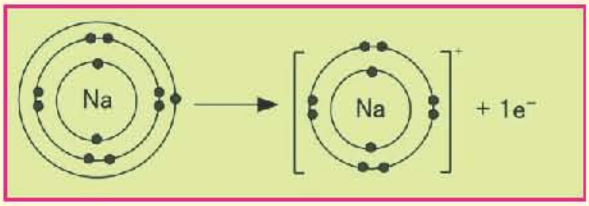
Example 5.1: Describing the formation of cations

Describe the formation of Na^{+} and Mg^{+2} cations.

Problem Solving Strategy:

1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. The sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.

2. Magnesium belongs to Group IIA in the periodic table. It has two valence electrons. A magnesium atom loses these electrons to achieve noble gas configuration. Represent this by drawing the complete electronic configuration or using an electron dot structure. This number also corresponds to the Group number in the periodic table.

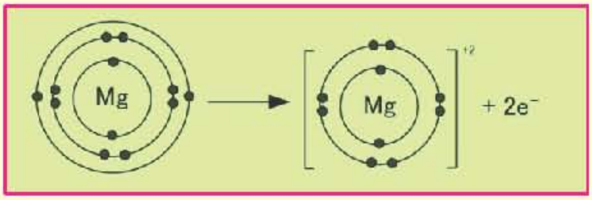


Solution:

(a) Formation of Na^{+} ion

$_{11}Na(1s^2 2s^2 2p^6 3s^1) \xrightarrow{-e^-} Na^{+} \ (1s^2 2s^2 2p^6)$

You can also represent this by following electron dot structure,



(b) Formation of Mg^{+2} ion

$$_{12}Mg\ (1s^2 2s^2 2p^6 3s^2) \xrightarrow{-2e^-} Mg^{2+}\ (1s^2 2s^2 2p^6)$$

You can also represent this by electron dot structure,

CONCEPT ASSESSMENT EXERCISE 5.1

1. Describe the formation of cations for the following metal atoms:
   (a) Li(atomic no 3)
   (b) Al(atomic no. 13)
2. Represent the formation of cations for the following metal atoms using electron dot structures.
   (a) K (b) Ca

Example 5.2: Describing the formation of anions.

Describe the formation of anions for the following non-metal atoms:

(a) Oxygen(atomic no.8) (b) Fluorine (atomic no. 9)

Problem Solving Strategy:

1. Write electronic configuration or dot structure.
2. Find the number of electrons needed to acquire eight electron configuration.
3. Represent addition of electrons.

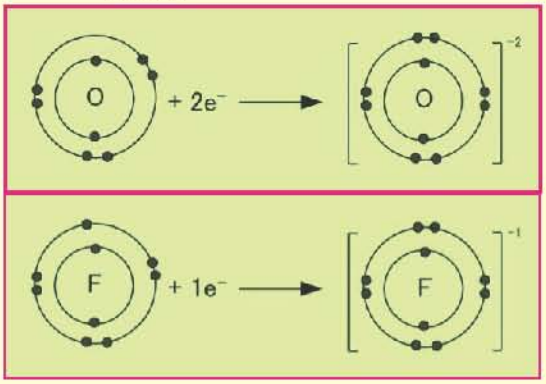
Solution:

(a) Formation of anion by oxygen atom.

Oxygen belongs to Group VIA on the periodic table. So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.

$$_8O(1s^2 2s^2 2p^4) + 2e^- \longrightarrow O^{2-}\ (1s^2\ \underbrace{2s^2 2p^6}_{\text{octet}})$$

You can also represent this by electron dot structure,



(a) Formation of anion by fluorine atom

Fluorine belongs to Group VIIA on the periodic table. So it has seven electrons in the valence shell. A fluorine atom therefore, requires only one electron to complete octet.

$$_9F(1s^2 2s^2 2p^5) + e^- \longrightarrow F^-\ (1s^2\ \underbrace{2s^2 2p^6}_{\text{octet}})$$

You can also represent this by electron dot structure,

CONCEPT ASSESSMENT EXERCISE 5.2

1. Describe the formation of anions by the following non-metals.
   (a) Sulphur (atomic No. 16) (b) Chlorine(atomic No. 17)
2. Represent the formation of anions by the following non-metals using electron dot structures.
   (a) N (b) P (c) Br (d) H
3. Compare differences between the formation of cations and anions.

Anions and cations have opposite charges. They attract one another by strong electrostatic forces. **"An ionic bond is a strong electrostatic attraction between positively charged metal ions and negatively charged non-metal ions"**. Compounds that consist of ions joined by electrostatic forces are called ionic compounds. The total positive charge of the cations must be equal to the total negative charge of the anions. This is because ionic compounds are electrically neutral as a whole.

Example 5.3: Representing ionic bond formation.

For each of the following pairs of atoms, use electron dot & electron cross structures to write the equation for the formation of ionic compound.

(a) Na and Cl (b) Mg and F

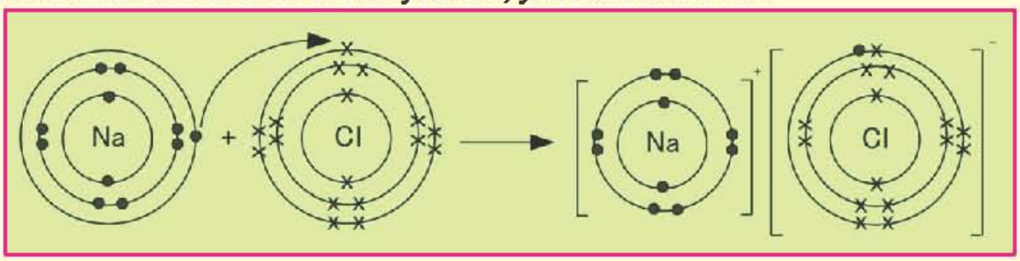
Problem Solving Strategy:

1. The metal atoms form cations and non-metal atoms form anions.
2. The number of electrons lost by metal atoms of group IA, IIA and IIIA equals the group number.
3. To write the final form of the equation, you need to know the simplest ratio of cations to anions that you require for the neutral compound.
4. Write equation using electron dot and electron cross structures.

Solution:

(a) Na is metal and Cl is non-metal.

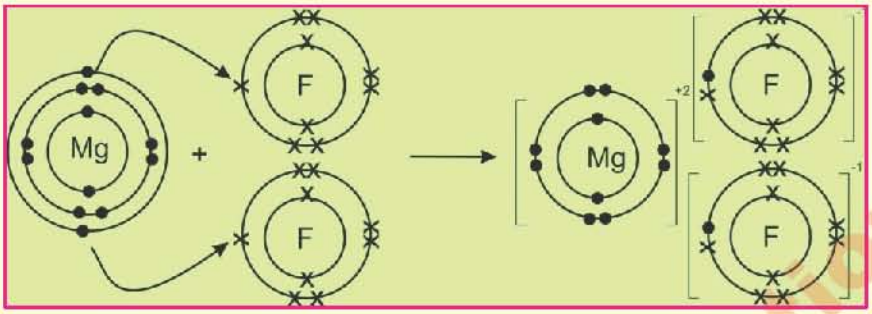
Metal atom tends to lose electrons and non - metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas. Since a Na atom has one electron in the outer most shell. It losses one electron to form Na^+ ion. Since a Cl atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form Cl^- ion. For every Na^+ ion, you need one Cl^- ion.



$2Na + Cl_2 \longrightarrow 2NaCl$ (Sodium Chloride)

(b) Mg is metal and F is non-metal.

A Mg atom has two electrons in the outermost shell. It losses two electrons to form Mg^{+2} ion. Since a F atom has seven electrons in the outermost shell, so it gains one electron to form F^- ion.



$$Mg + F_2 \longrightarrow MgF_2 \text{ (Magnesium Fluoride)}$$

For every Mg^{+2} ion you need two F^- ions.

CONCEPT ASSESSMENT EXERCISE 5.3

For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound.

(a) Mg and O (b) Al and Cl

Example 5.4: Recognizing a compound as having ionic bonds.

Recognize the following compounds as having ionic bonds.

(a) MgO (b) NaF

Problem Solving Strategy:

1. The metal atom loses electrons to form cations and non-metal atom gains electrons to form anions.
2. The number of electrons lost by metal atoms of group IA, IIA and IIIA equals the group number. The number of electrons gained by the non-metal atoms is equal to 8 minus group number.
3. Find the simplest ratio of cations to anions, to identify the compound.

Solution:

(a) MgO

Mg is metal and O is non-metal. A Mg atom has two electrons in outermost shell. So it loses two electrons to form Mg^{+2} ion. Since an O atom has six electrons in outermost shell, so it gains two electrons to form O^{-2} ion. In this way both the atoms acquire nearest noble gas configuration. For every Mg^{+2} ion you need one O^{-2} ion. Chemical formula of resulting compound is MgO. Therefore MgO is an ionic compound.

(b) Na is metal and F is non-metal. A Na atom has one electron in outmost shell. So it loses one electron to form Na^+ ion. Since a F atom has seven electrons in outermost shell, so it gains one electron to form F^- ion. Na atom by losing one electron and F atom by gaining one electron acquire nearest noble gas electronic configuration. You need one F^- ion for each Na^+ ion. Therefore, NaF is an ionic compound.

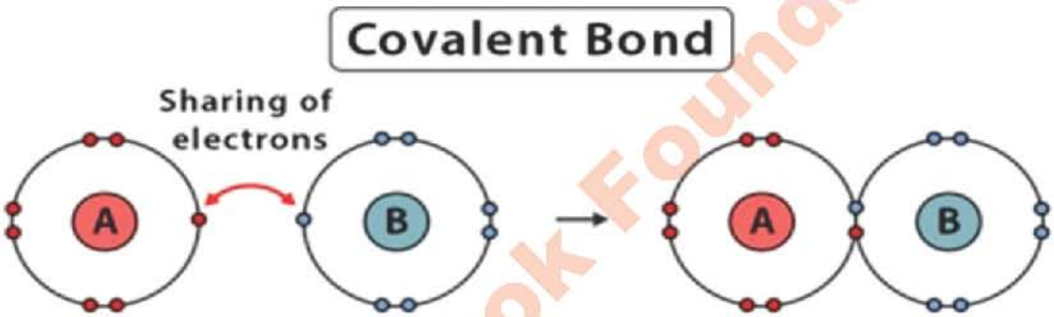
CONCEPT ASSESSMENT EXERCISE 5.4

Recognize the following compounds as having ionic bonds:

(a) KCl (b) $AlCl_3$ (c) MgF_2 (d) NaF (e) NaBr

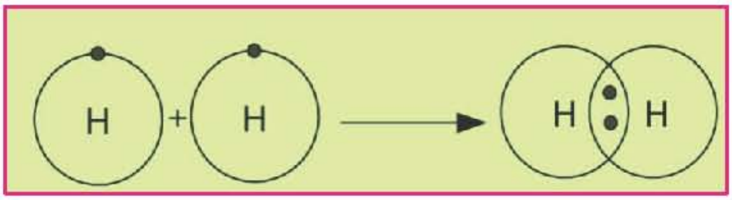
5.3.2 Covalent Bonds

Nonmetal atoms tend to share electrons with each other or with other nonmetal atoms, forming a chemical bond called a covalent bond. A chemical bond formed by mutual sharing of electrons between two atoms is called a covalent bond. General representation of a covalent bond is given below.



Consider the formation of a covalent bond between two hydrogen atoms. A hydrogen atom has one valence electron. Two hydrogen atoms share their valence electrons to form a diatomic molecule.

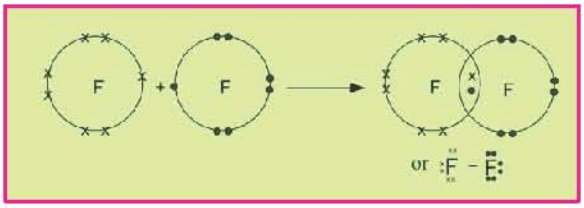
$H\cdot + \cdot H \longrightarrow H:H$ shared pair of electrons



In the formation of this molecule, each hydrogen atom reaches the electronic configuration of the noble gas helium with two valence electrons. An electron pair in the region between two atoms attracts both hydrogen nuclei. This creates a strong electrostatic attraction between the shared electrons and the two nuclei. This means that the situation is more stable than in individual atoms. Because of this stability, the two atoms form a covalent bond.

In a covalent bond, a strong electrostatic forve of attraction between the bonding electrons and two atomic nuclei binds them together.

A covalent bond between two atoms can be represented by using electron-dot and electron-cross symbols for the atoms and the resulting molecule. As already discussed valence electrons are represented by dots. Just to understand sharing, we represent valence electrons in one atom by dots and in the other atom by crosses. However, remember that all the electrons are identical and

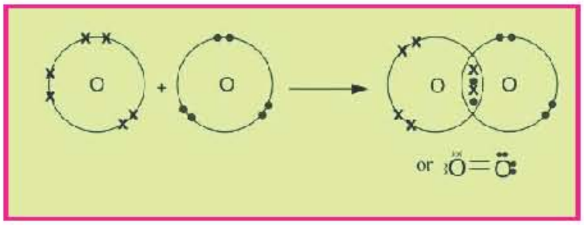


cannot be differentiated. A shared pair of electrons is also represented by a dash (-) in a molecule.

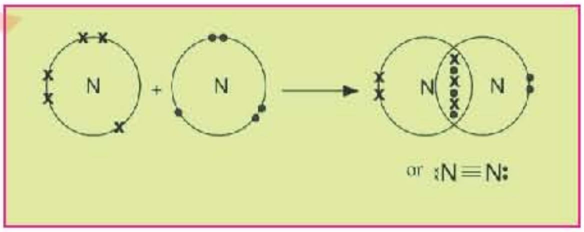
Consider the formation of a bond between two fluorine atoms. Fluorine belongs to Group VIIA, so it has seven electrons in the valence shell. It needs one more electron to attain the electron configuration of a noble gas. Thus two F-atoms share an electron pair and achieve electron configuration of Ne. For sharing each F-atom contributes one electron to complete the octet.

Pairs of valence electrons that are not shared between atoms are called **lone** pairs or lone pairs. A covalent bond formed by sharing **one pair of electrons** is called a single covalent **bond**. So **both H_2 and F_2** molecules contain **single** covalent bond.

Can you explain the formation of covalent bond between H-atom and a F-atom?



Sometimes atoms may share two or three electron pairs to complete an octet. **Double covalent bonds are the bonds that are formed by sharing of two electron pairs. Triple covalent bonds are the bonds that involve three shared pairs of electrons.**



Consider the formation of O_2 molecules. Oxygen is in Group VI A, so it has 6 electrons in the valence shell. It needs two electrons to complete its octet. So for sharing each O-atom contributes two electrons.

Can you explain the formation of N_2 molecules?

Example 5.5: Drawing electron cross and dot structures for simple covalent molecules containing single covalent bonds

Draw electron cross and dot structures for (a) CH_4 that is a major component of natural gas (b) H_2O that covers about 80% of the earth crust.

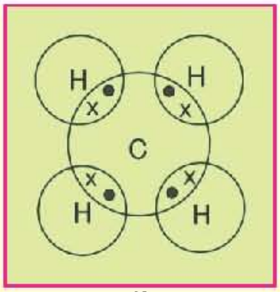
Problem Solving Strategy:

1. Decide from the chemical formula which atom is the central atom. An atom that contributes more electrons for sharing is the central atom. Show its valence electrons by dots. Note the number of electrons it needs to complete octet. If the number of electrons needed equals the other atoms, each atom will form a single covalent bond.
2. Arrange other atoms around the central atom. Connect the central atom by single bonds. Use cross to represent electrons of the other atoms.
3. Check whether the arrangement of electron satisfies the octet rule.

Solution:

(a) CH_4

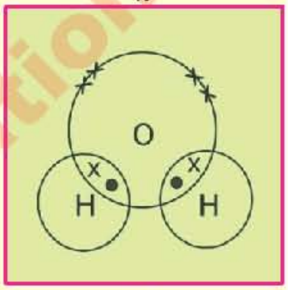
(i) C has four electrons in the valence shell and needs four electrons to complete its octet. H has only one valence electron and needs one electron to complete the duplet. So C can form four single bonds with four H-atoms. C is the central element.



(ii) Connect the atoms with a dot and a cross

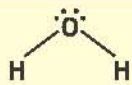
(b) H_2O

(i) O has six valence electrons ·Ö· and each hydrogen atom has one valence electron. Ḣ So O-atom needs two electrons to complete the octet. Each H needs one electron to complete duplet.



(ii) O is central atom and will form two single bonds with H-atoms.

(iii) Arrange H-atoms around O and connect them by a pair of electrons (one dot and one cross)



CONCEPT ASSESSMENT EXERCISE 5.5

Draw electron cross and dot structures for the following molecules:

(a) NH_3

(b) HCl

(c) CH_3OH

Example 5.6: Drawing electron cross and dot structures for molecules containing multiple bonds

Draw electron cross and dot structures for the following molecules:

(a) CO_2, a component of air and is responsible for greenhouse effect.

(b) HCN, used as insecticide.

Problem Solving Strategy:

1. Decide from the formula which atom is to be in the center. Show its valence electrons by dots. Note the number of electrons it needs to complete octet.
2. Show valence electron of the other atoms by cross and find the number of electrons each of the atoms needs to complete octet or duplet.
3. Connect central atom with the other atoms by electron pair or pairs to satisfy the octet rule.

Solution:

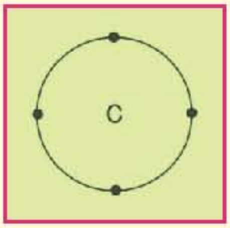
(a) CO_2

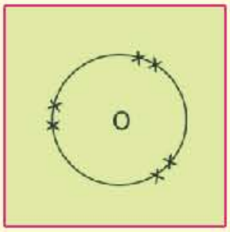
(i) C has four electrons in the valence shell. It needs four electrons to complete octet.

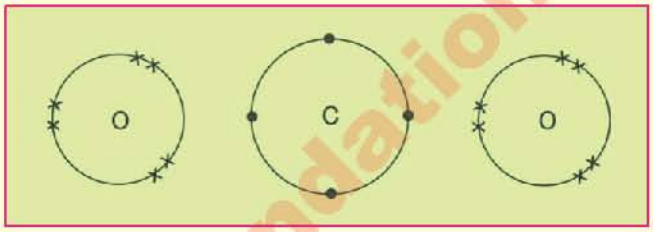
(ii) Each oxygen atom has six valence electrons and needs two electrons to have an octet.

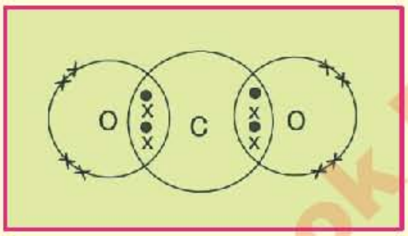
(iii) C is central atom, arrange O-atoms around it.

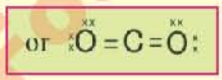
(iv) Since C needs four electrons and there are only two oxygen atoms. So it will share its two electrons with each oxygen atom.







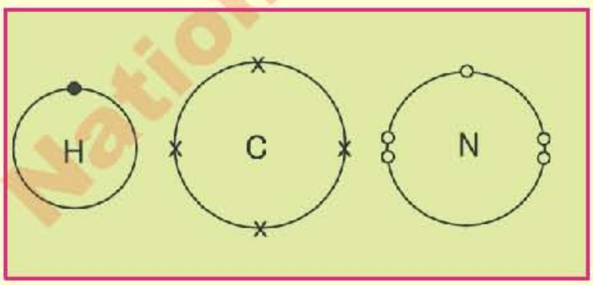


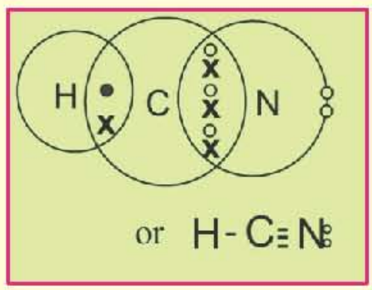


(b) HCN

(i) H has one, C has four and N has five electrons.

(ii) C needs four and N needs three electrons. So C shares one electron with H to form a single bond and three electrons with N to form a triple bond. This will satisfy octet rule.





CONCEPT ASSESSMENT EXERCISE 5.6

Draw electron cross and electron dot structures for the following molecules:

(a) CS_2 an organic solvent that dissolves sulphur, phosphorus etc

(b) N_2 a component of air.

(c) C_2H_4, ethane, a component of natural gas.

5.3.3 Types of covalent bond on the basis of polarity:

Non-Polar Covalent bond:

A covalent bond can **form** between two similar atoms such as in H_2, N_2, O_2, Cl_2 etc. It can also occur between two **different** atoms, as in, HCl, H_2O, NH_3, HCN, CO_2etc. When two identical atoms share electron pairs, both atoms exert the same force on the shared electron pairs. Such a covalent bond is called a **nonpolar** covalent bond. For example, bonds H-H, O = O, etc. are non-polar covalent bonds.

Polar Covalent bond:

On the other hand, when two different atoms share an electron pair, both atoms exert different forces on the shared electron pair. **The more** electronegative atom pulls **the** shared electron pairs towards **itself** with **a** greater force than the **other atom. Thus, the** more electronegative atom **attracts some of the** electron density towards itself. This makes it partially negatively charged and **the** other atoms partially positively charged. Such a covalent bond is called a polar covalent bond. The forces of attraction between molecules are called intermolecular forces. For example, $\overset{\delta+}{H}-\overset{\delta-}{Cl}$:

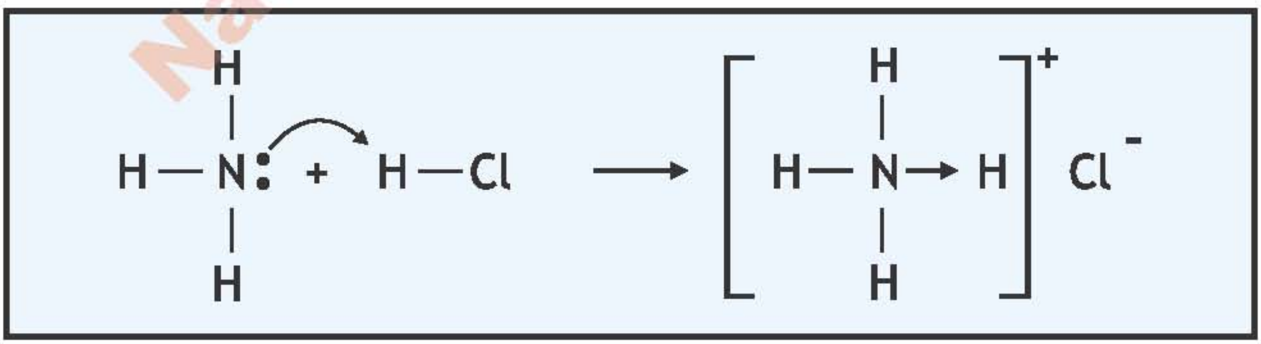
5.3.4 Coordinate Covalent Bond

A coordinate covalent bond is a type of covalent bond where the shared electron pair comes from a single atom (called donor). Atoms are held together because both nuclei attract a pair of electrons. Once a covalent bond is formed, it is impossible to distinguish the origin of the electrons. Such bonding is usually observed when metal ions bind to ligands. However, nonmetals can also participate in this bond. The reaction between a Lewis acid and a base is a covalent coordinate bond.

Examples of coordinate covalent bonds:

1. **Ammonium (NH_4^+) ion**

   The ammonium ion is formed from the reaction of ammonia (NH_3) gas with hydrogen chloride (HCl) gas. In NH_4+, the fourth hydrogen is attached by acoordinate covalent bond because only the hydrogen's nucleus is transferred from the chlorine to the nitrogen. The



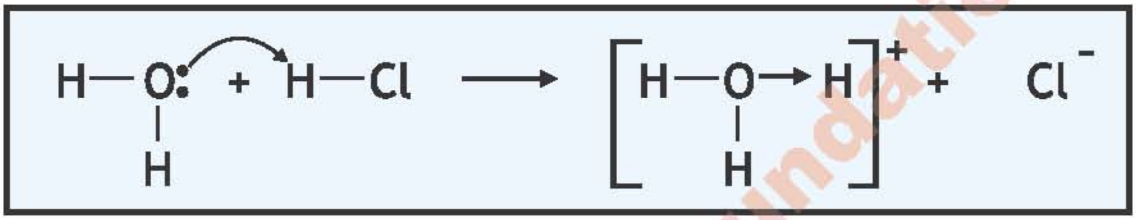
hydrogen's electron is left behind on the chlorine to form a negative chloride (Cl^-) ion.

2. **Hydronium ion (H_3O^+)**

When hydrogen chloride (HCl) gas dissolves in water to make hydrochloric acid (HCl aq.), a coordinate covalent bond is formed in the hydronium ion. The hydrogen (H) nucleus is transferred to the water (H_2O) molecule, which has a lone pair of electrons to form hydronium. So, H does not contribute any electrons to the bond.

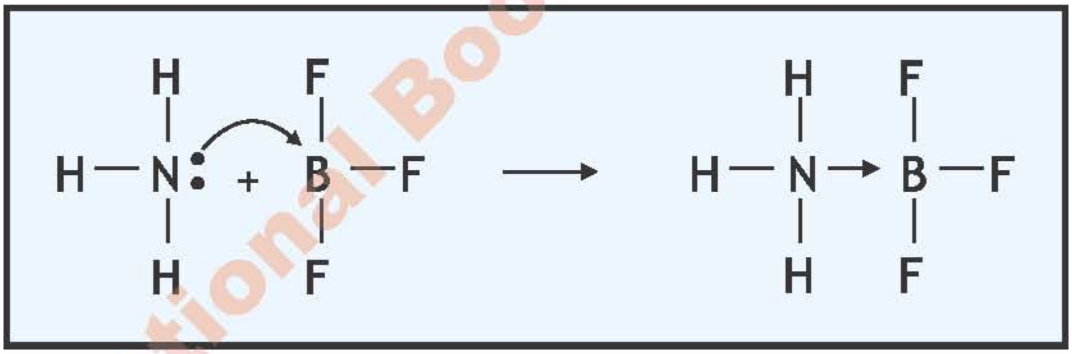
3. **Ammonia Boron Trifluoride (NH_3-BF_3)**

Boron trifluoride (BF_3) is a compound that does not have a noble gas structure around the boron (B) atom. The boron only has three pairs of electrons in its valence shell and requires a pair to complete the orbital. Hence, BF_3 is electron deficient. The lone pair on the nitrogen (N) of the ammonia (NH_3) molecule is used to overcome that deficiency, and a



CONCEPT ASSESSMENT EXERCISE 5.7

1. Differentiate between polar and non-polar covalent bonds.
2. How is coordinate covalent bond different from normal covalent bond?



5.4 INTERMOLECULAR FORCES

An intermolecular force is the attractive force that exist between the molecules.

Dipole-dipole forces

Dipole-dipole interactions occur between polar molecules. Figure 5.1 shows these interactions.

You know that paints and dyes are used to protect solid surfaces from the atmospheric effects. They also give visual appeal. Resins are used to coat materials that give toughness, flexibility, adhesion and chemical

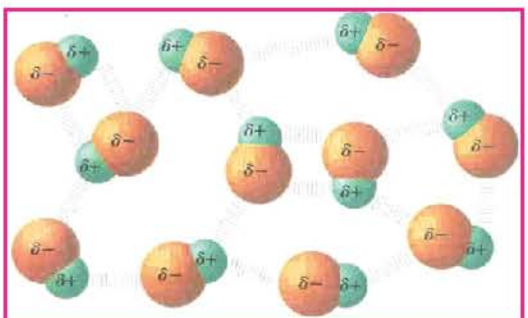


Figure 5.1: Dipole-Dipole interactions

resistance. For example dams, bridges, floors, trains, buses, cars etc are painted with resins. The synthetic resins are used where water resistance is required. Chemically, resins are either adhesive or they form bond linkages with the material being bonded together. What is the nature of these linkages?

Notice that slightly negative end of polar molecule is weakly attracted to the slightly positive end of another molecule. Such attracting forces are called dipole-dipole interactions.

Hydronging bonding

Molecules in which hydrogen is covalently bonded to a very electronegative atom such as oxygen, nitrogen or fluorine is also weakly bonded to a lone pair of electron of another electronegative atom. This other atom may occur in the same molecule or in a nearby molecule. This intermolecular interaction is called hydrogen bonding. Oxygen, nitrogen or fluorine makes hydrogen very electron-deficient. Thus interaction of such a highly electron deficient hydrogen and lone pair on a nearby electronegative atom compensates for the deficiency. Figure 5.2 shows hydrogen bonding in water molecules.

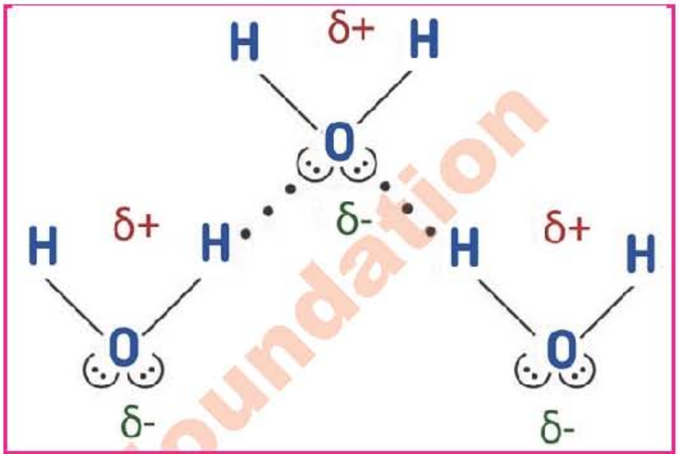


Figure 5.2: Hydrogen bonding in water

The interaction of a highly electron deficient hydrogen and lone pair on a nearby highly electronegative atom such as N, O or F is called hydrogen bond. This phenomenon is called hydrogen bonding.

These intermolecular forces are extremely important in determining properties of water, biological molecules, such as proteins, DNA etc and synthetic materials such as glue, paints, resins etc. The adhesive action of paints and dyes is developed due to hydrogen bonding. Synthetic resins bind two surfaces together by hydrogen bonding or dipole-dipole interactions

Society, Technology and Science

Epoxy adhesives have excellent chemical resistance, good adhesion properties, good heat resistance and they form strong and tough coating. Therefore, propellers and parts of aircraft, boats, cars, trucks etc are held together by epoxy adhesives. Epoxy adhesives contain partially positively charged H-atoms and oxygen atoms containing lone pairs in their molecules. Epoxy adhesives are, therefore, sticky and can make H-bonds with other substances. Modern aircraft, boats and automobiles such as cars, trucks etc and even in space craft epoxy adhesives are used for assembling, saving money and reducing weight. This means glues and adhesives have become an essential item in our daily life.

5.5 NATURE OF BONDING, STRUCTURE AND PROPERTIES

Three main factors are important when determining the properties of a substance:

1. **Type of Particles**

   The types of elementary particles contained. The substance can contain atoms, ions or molecules. For example, if it contains ions (such as sodium chloride), it will conduct electricity when melted or dissolved in water. In order to be soluble in water, the substance must contain ions or polar molecules.

2. **The way elementary particles are connected to each other.**

Particles may have ionic, covalent, metallic, or weak intermolecular forces. The stronger the bond, the higher the melting/boiling point and hardness of the substance.

For example, silicon dioxide (SiO_2) has strong covalent bonds, connecting each atom to several other atoms to form a giant covalent structure. The atoms in silica are difficult to separate, making it very hard and difficult to melt.

On the other hand carbon dioxide has strong covalent bonds between the C and O atoms. But these molecules have weak intermolecular forces between them. The molecules are therefore easily separated and so CO_2 has a low melting/boiling point.

3. **The arrangement of particles**

Particles may be arranged in planes (for example, polymers), in layers (for example, clays, graphite) or in a variety of three-dimensional networks. In graphite atoms are arranged in 2-dimensional layers. This allows the layers of graphite to move over one another .(for example, graphite pencil writing). Diamonds have a large three-dimensional network of carbon atoms, which make it the hardest substance on earth. Metals also have giant structures. metallic bonding is stong , most metals have very high melting and boiling points and are thermally stable.

Conduction of electricity in ionic compounds

Electrical conductivity is achieved by the movement of charged particles. Ionic compounds cannot conduct electricity in the solid state because their ions remain in a fixed position and cannot move. When an ionic compound is melted or dissolved in water. It is ionized, its ions move freely in molten or aqueous solution. Therefore electricity can pass through a molten ionic compound or its aqueous solution.

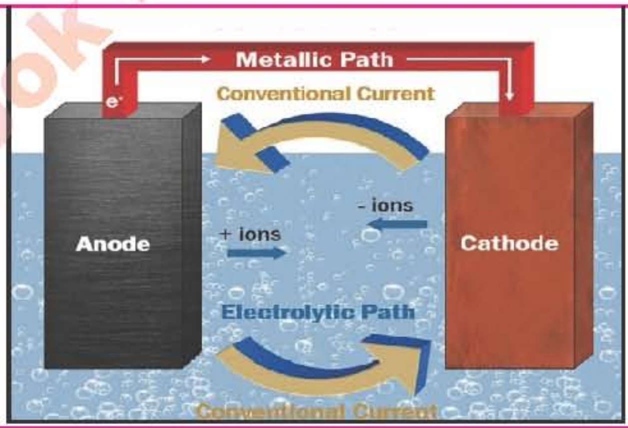


Figure 5.4: conduction of electricity through molten NaCl

Conduction of electricity through acids

Covalent compounds have no free charged particles, so they do not conduct electricity. However, some covalent compounds conduct electricity when dissolved in water. For instance, acids like HCl, H_2SO_4, HNO_3 etc. When these acids are dissolved in water, they ionize and form high concentrations of H+ ions and negatively charged ions. These ions can move freely in aqueous solution. Therefore, aqueous solutions of acids conduct electricity.

Metals are good conductor of electricity because they have free electrons. These electrons are not associated with a single atom. These electrons begin to flow under the influence of electricity. Therefore metals allow electricity to pass through.

Compounds that consists of covalent molecules are called covalent compounds. The intermolecular forces between their molecules are much weaker than the covalent bonds. Therefore, covalent compounds have low melting and boiling points. Since their molecules do not contain any free electrons or ions, they are poor conductors of electricity.

Intermolecular Forces and Their Influence on the Melting and Boiling Points

Tables shows melting and boiling points of some common covalent and ionic compounds.

Table 5.1: Melting point and boiling points of some covalent compounds

Compound	Melting Point (^{0}C)	Boiling Point (^{0}C)
Water (H_2O)	0	100
Mehtane (CH_4)	-183	-162
Ethanol (CH_3CH_2OH)	-117	78

Table 5.2: Melting point and boiling points of some ionic compounds

Compound	Melting Point (^{0}C)	Boiling Point (^{0}C)
Sodium Chloride (NaCl)	801	1465
Sodium Fluoride (NaF)	996	1695
Magnesium Chloride ($MgCl_2$)	714	1412

Covalent compounds usually have much lower melting points than ionic compounds. For example, a common covalent compound of water has a melting point of 0°C and a boiling point of 100°C. The melting points and boiling points of the common ionic compound sodium chloride are 801°C and 1465°C. This is because ionic compounds involve breaking the ionic bond. Breaking the electrostatic forces between ions requires large amounts of energy. Thus, ionic compounds have high melting points and boiling points. Melting of covalent solids involves the breaking of intermolecular forces, which are much weaker than electrostatic forces. Thus, less energy is required to break the intermolecular forces between covalent molecule

5.5.1 Graphite

Graphite's name is derived from the Greek word "graphein," meaning "to write." It is commonly called black lead. Graphite is an allotrope of carbon. Graphite is formed when carbon is subjected to the intense heat and pressure of the earth's crust and upper mantle.

Structure of Graphite

In graphite, each carbon atom is linked with 3 other carbon atoms by a single covalent bond resulting in the hexagonal ring arranged in a layer. It has a 2-dimensional layers structure. The 4th valence of the carbon atom is satisfied by weak Vander walls forces between 2 layers.

Uses of graphite

1. Graphite is a unique material since it has both metal and non-metal qualities. Moreover, it is a soft mineral with black colour, slippery surface and lustre. These properties are due to layered structure of graphite. Its major uses include:

2. Due to its stability in high temperatures and chemical inertness, graphite is used in many refractory items such as carbon refractory bricks.
3. The electrodes of graphite are used in electrical metallurgical furnaces. It is used as an anode in electrolytic processes.
4. Graphite is used in making moderator rods and reflector components in a nuclear reactor. It is used in the manufacturing of carbon brushes andelectric motors.
5. Graphite material is used in engineering sectors in the making of thrust and journal bearing, piston rings, and vanes.
6. Other applications of graphite include metallurgy, as lubricants, and in the production of paints and pencils.

All these uses are a testament to the unique properties of graphite. The patterned bonding and layered structure make it suitable for such diverse applications.

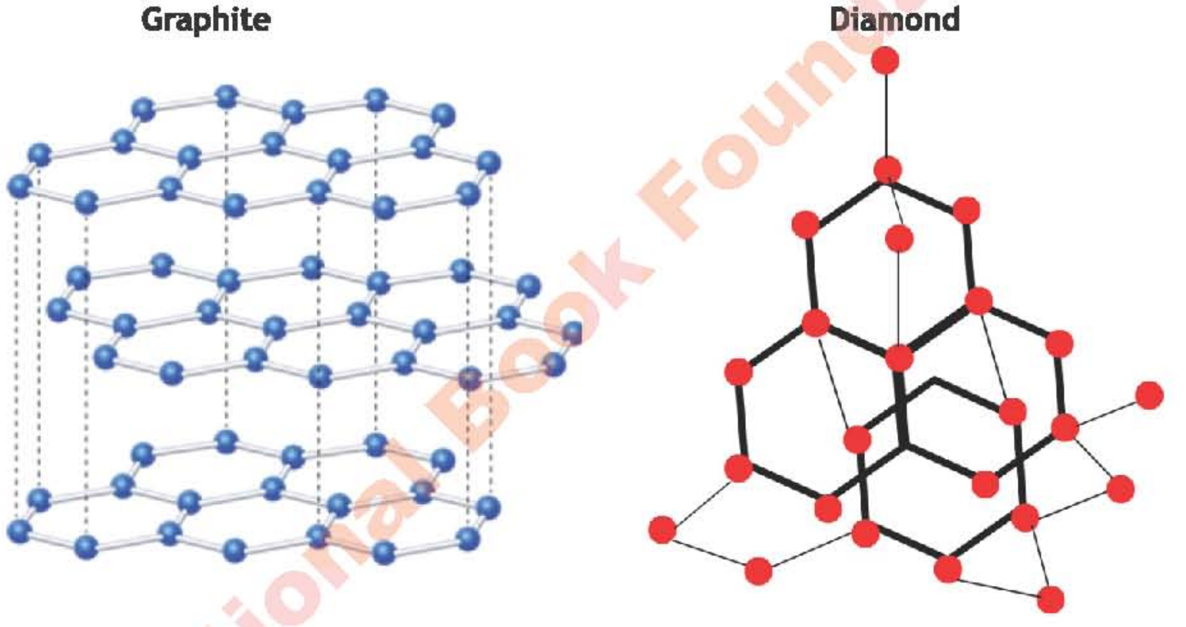


Figure 5.5: Structure of graphite and dimond

5.5.2 Diamond

Diamond is an allotrope of carbon in which the carbon atoms are arranged in a diamond cubic crystal lattice. Thanks to the presence of **strong covalent bonds** and a rigid tetrahedral structure, Diamond is the **hardest** material ever discovered.

Structure of Diamond

In a diamond, the carbon atoms are arranged tetrahedrally. Each carbon atom is attached to four other carbon atoms 1.544×10^{-10} meter away with a C-C-C bond angle of 109.5°. It is a strong, rigid three-dimensional structure that results in an infinite network of atoms. This accounts for diamond's hardness, extraordinary strength and durability and gives diamond a higher density than graphite (3.514 grams per cubic centimeter).

Properties and uses of Diamond

The giant structure and extensive covalent bonding in diamond renders it extraordinary hardness, elasticity, high yield strength, less conductivity, and chemical inertness. Owing to these properties diamond has variety of applications like:

1. Diamonds are most commonly used in ornaments like rings, necklace, earrings, etc. In the gem industry, the value of diamonds is very high. They are used in making jewellery because of their durability and lustre property.
2. Its property of hardness is useful to drill, grind or cut materials. Hence, some blades used for cutting and drills in the industry used diamonds. They are present on the edges and tips in small sizes.
3. Diamonds are used in making medicines and beauty products. They are also used in making medical tools, like tools used in cataract surgery. Nano-diamonds have potential health benefits.
4. Diamonds produce high-quality sound because they are hard and vibrate easily at high speed. It is also used in DJ equipment and high-quality recorders.

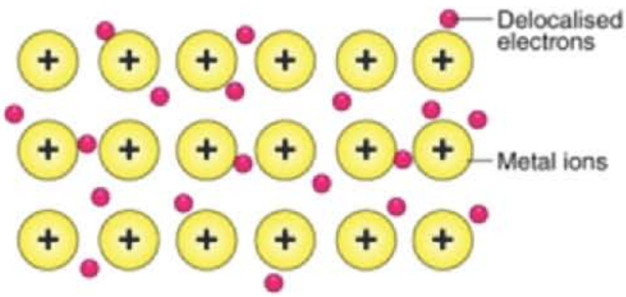
5.5.3 Contrasting ionic and covalent compounds and their uses

The type of chemical bonds significantly influences the properties and uses of materials.

- Ionic compounds are strong in compression, but they are brittle, i.e. they can break easily. In the solid state ionic compounds are poor conductors of electricity. But when they melt or dissolve in water, they conduct electricity due to the free movement of ions. Therefore, batteries and fuel cells use ionic compounds as electrolytes.
- Covalent compounds with giant structures, such as diamond, quartz, silica, etc. are usually very strong and hard. Because of its hardness, diamond is used in cutting and drilling tools. Quartz and silicon dioxide are used in the production of abrasives. Graphite, quartz and silica, because they are stable at high temperatures, are used to make ceramics, glass and refractories. Most covalent compounds are poor conductors of electricity

5.6 METALLIC BONDS

A special type of bonding occurs in metals. In metals, the valence electrons are not confined to individual atoms. These electrons are called free electrons. Metal atoms lose these electrons and form positive ions. The free electrons can move throughout the entire metal structure. This leads to the forming a sea of delocalized electrons called the electron sea. The metal cations are held together by the strong electrostatic attractive forces between the metal cations and negatively charged electron sea. This force gives metals their unique properties. This type of bonding is called metallic bonding.



The properties of metals that are a consequence of metallic bonding include:

Malleability

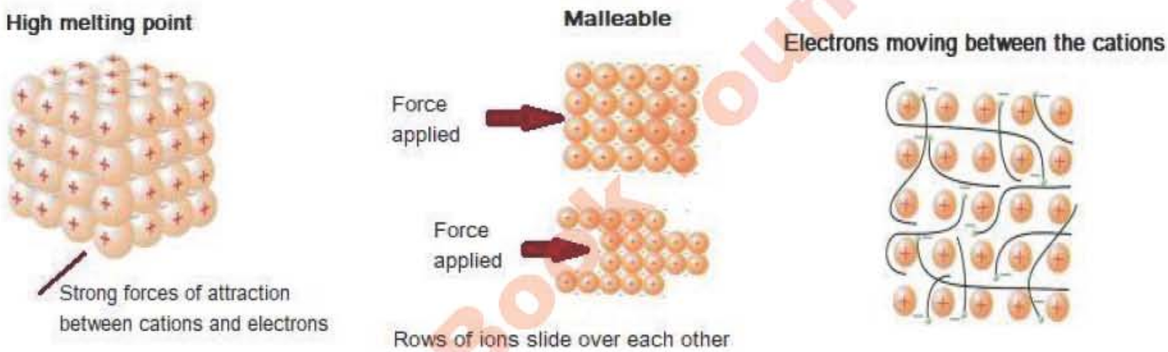
Ductility

High melting and boiling point

High electrical and thermal conductivity Metallic lustre

5.6.1 Structure and Properties of Metals Which make it Suitable for Industrial Purposes

1. Metals have giant structures. Metallic bond is strong due to which metals have very high melting and boiling points. This makes them thermally stable.
2. The layers are able to slide over each other, which makes the metals to bent and shaped. This makes them malleable and ductile. They can be drawn into wires and sheets.
3. Metals are good conductors of electricity because the delocalised electrons can move freely. The delocalised electrons can also transfer energy from one place to another and conduct thermal energy.



KEY POINTS

- An octet is a set of eight. In order to gain stability atoms tend to gain electron configuration of nearest noble gas.
- The tendency of atoms to acquire eight electron configuration in their valence shell, when binding is called octet rule.
- Ionic bonds are formed between two atoms, when one atom loses electrons and other atom gains these electrons. The force of attraction that binds oppositely charged ions is called ionic bonds.
- Ionic compounds have high melting points. They conduct electricity in molten state.
- A bond that is formed by the sharing of electrons between two atoms is called a covalent bond. A covalent bond can be single, double or triple.
- The interaction of a highly electron deficient hydrogen and lone pair on a nearby electronegative atom is called hydrogen-bond.
- The adhesive action of paints and dyes is developed due to hydrogen bonding.

References for additional information

- Lawarie Ryan, Chemistry for you.
- Iain Brand and Richard Grime, Chemistry (11-14).
- Silberg, Chemistry.
- Raymond Chang, Essential Chemistry.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which of the following atoms will form an ion of charge -2?

	Atomic Number	Mass Number		Atomic Number	Mass Number
(a)	12	24	(b)	14	28
(c)	8	8	(d)	10	20

(ii) Which of the following atoms will not form cation or anion.

(a) (Atomic No. 16) (b) (Atomic No. 17)

(c) (Atomic No. 18) (d) (Atomic No. 19)

(iii) Which of the following atoms will form cation.

	Atomic Number		Atomic Number
(a)	20	(b)	18
(c)	17	(d)	15

(iv) Which of the following atoms obey duplet rule?

(a) O_2 (b) F_2

(c) F_2 (d) N_2

(v) Silicon belongs to Group IVA. It has ____ electrons in the valence shell

(a) 2 (b) 3

(c) 4 (d) 6

(vi) Phosphorus belongs to third period of Group VA. How many electrons it needs to complete its valence shell.

(a) 2 (b) 3

(c) 4 (d) 5

(vii) In the formation of AlF_3, aluminum atom loses _____ electrons.

(a) 1 (b) 2

(c) 3 (d) 4

(viii) Which of the following is not true about the formation of Na_2S:

(a) Each sodium atom loses one electron

(b) Sodium forms cation

(c) Sulphur forms anion

(d) Each sulphur atom gains one electron

(ix) Identify the covalent compound

(a) NaCl (b) MgO

(c) H_2O (d) KF

2. Give short answer.

(i) State octet and duplet rules.

(ii) Explain formation of covalent bond between two nitrogen atoms

(iii) How does Al form cation?

(iv) How does O from anion?

(v) Draw electron cross and dot structure for H_2O molecule.

3. Describe the importance of noble gas electronic configuration.

4. Explain how elements attain stability?

5. Describe the ways in which bonds may be formed.

6. Describe the formation of covalent bond between two non-metallic elements.

7. Explain with examples single, double and triple covalent bond.

8. Find the number of valence electrons in the following atoms using the periodic table:

(a) Boron (b) Neon

(c) Rubidium (d) Barium

(e) Arsenic

9. Represent the formation of cations for the following metal atoms using electron dot structures.

(a) Al (b) Sr (c) Ba

10. A sample of sulphur from a volcano was analysed to give the following composition of isotops (At no of S = 16)

Isotope	Abundance (%)
S - 32	95.0
S - 33	0.76
S - 34	4.22

(a) Define the term isotope

(b) Define the term relative atomic mass

(c) Calculate the relative atomic mass of sulphur

(d) Complete the following table.

	Protons	Neutrons	Electrons
S - 32			
S - 34			

(e) Where will you place S in the periodic table.

(f) How many electrons S will loss or gain to acquire stable configuration.

(g) How many atoms of S are there in 0.3 mole of sulphur.

11. An atom of an element has atomic number 9 and mass number 19.

(a) State the number of protons and neutrons in the nucleus of this atom.

(b) State the number of electrons in this atom.

(c) Show with electron cross-dot diagrams, the formation of ions by this atom.

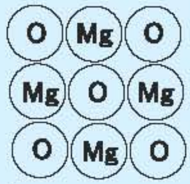
(d) Write electronic configuration of this element.

(e) Point out its group in the periodic table.

(f) Point out its period in the periodic table.

THINK TANK

12. Magnesium oxide is a compound made up of magnesium ions and oxide ions.

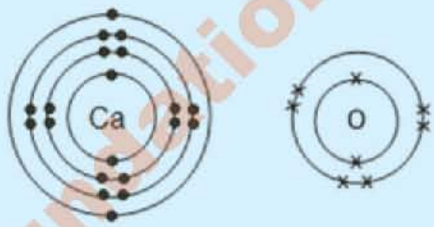


(a) What is the charge on these ions.

(b) How these ions get these charges.

(c) Show with electron cross-dot diagrams the formation of these ions.

13. The diagrams below show the electronic structures of an atom of calcium and an atom of oxygen.



Draw structures of the ions that are formed when these atoms react.

14. The table below shows the properties of four substances:

Substance	Melting point	Electrical Conductivity	
		In solid state	In molten state
A	High	NIL	NIL
B	High	NIL	Good
C	Low	NIL	NIL
D	High	Good	Good

(a) Which substance is a metal?

(b) Which substance is an ionic compound?

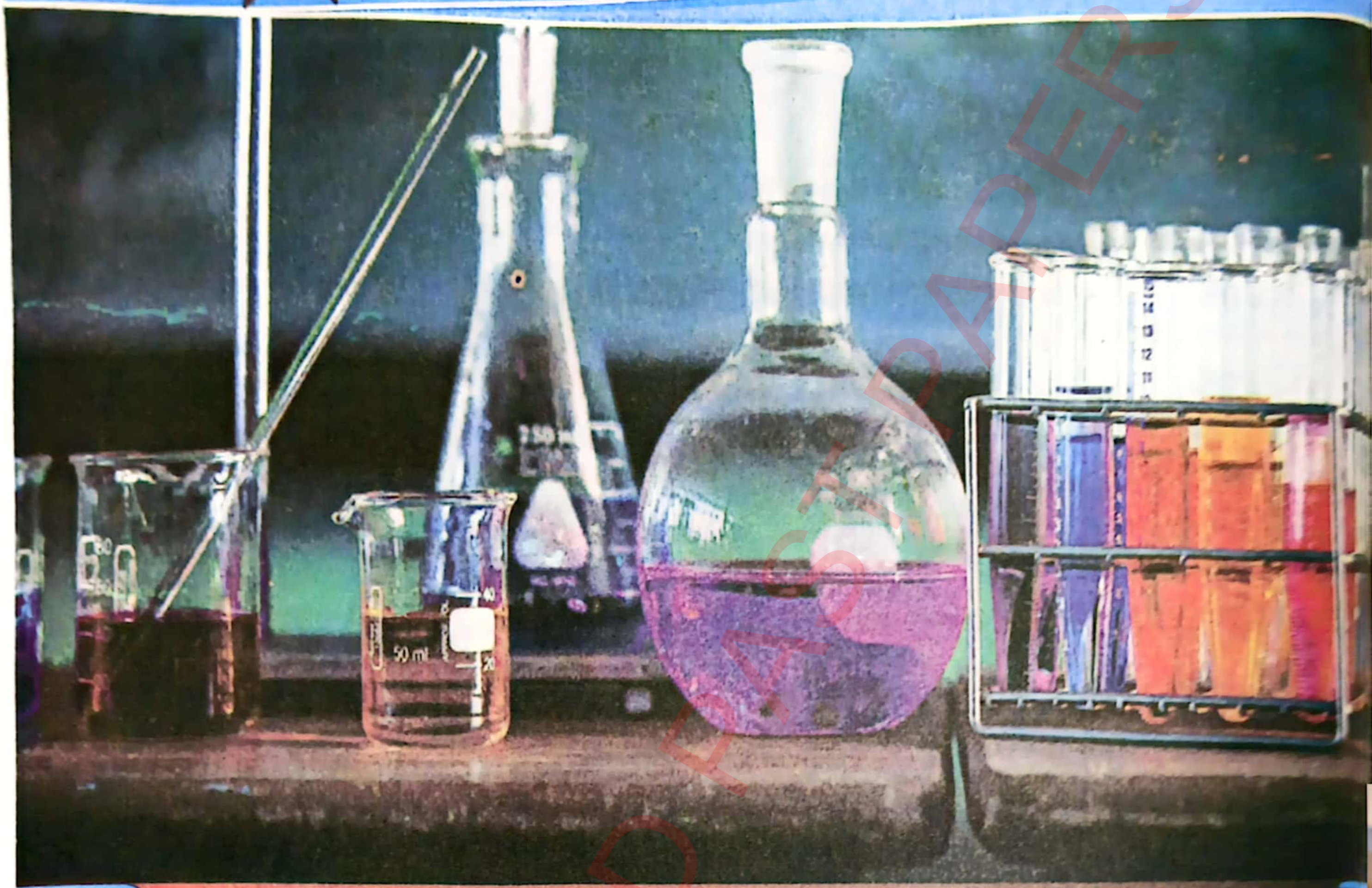
(c) Which substance is a covalent compound?

(d) Which substance is a non-metal?

PROJECT

Prepare a chart displaying different types of bonds with example.

UNIT 06



STOICHIOMETERY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- State the formulae of common elements and compounds.
- Define molecular formula of a compound as the number and type of different atoms in one molecule.
- Define empirical formula of a compound as the simplest whole number ratio of different atoms in a molecule.
- Deduce the formula and name of a binary compound from ions given relevant information.
- Deduce the formula of a molecular substance from a given structure of molecules.
- Use the relationship amount of substance = mass/molar mass to calculate number of moles, mass, molar mass, relative mass (atomic/molecular/formula) and number of particles.
- Define mole as amount of substance containing Avogadro`s number (6.022×10^{23}) of particles.
- Explain the relationship between a mole and Avogadro`s constant.
- Construct chemical equations and ionic equations to show reactants forming products, including state symbols.
- Deduce the symbol equation with state symbols for a chemical reaction given relevant information.

INTRODUCTION

What are the simplest components of wood, rocks and living organisms? This is an age-old question. Ancient Greek Philosophers believed that everything was made of an elemental substance. Some believed that substance to be water, other thought it was air. Some other believed that there were four elemental substances.

As 19th century began, John Dalton proposed an atomic theory. This theory led to rapid progress in chemistry. By the end of the century however, further observations exposed the need for a different atomic theory. 20th century led to a picture of an atom with a complex internal structure.

A major goal of this chapter is to acquaint you with the fundamental concepts about matter. In this chapter you will learn some basic definitions to understand matter. This knowledge will help you in grade XI.

6.1 EMPIRICAL FORMULA AND MOLECULAR FORMULA

Remember that the chemical formula of a compound tells you which elements it contains and the whole number ratio of those atoms. In a chemical formula, the elemental symbol and numerical subscripts indicate the type and number of each atom in the compound. The compound has several chemical formulas. Learn about the two types of chemical formulas.

6.1.1 Empirical Formula

The empirical formula of a compound is the chemical formula that gives the simplest integer ratio of the atoms of each element. For example, the compound hydrogen peroxide has one H atom for every O atom. Therefore, the simplest ratio of hydrogen to oxygen is 1:1. The empirical formula for hydrogen peroxide is therefore written as HO.

The simplest ratio between C, H and O atoms in glucose is 1 : 2 : 1. What is the empirical formula of glucose?

6.1.2 Molecular Formula

A molecular formula is an expression that specifies the number of atoms of each element in one molecule of a compound. This shows the actual number of atoms in each molecule. For example, the molecular formula of hydrogen peroxide is H_2O_2. It shows that each molecule of hydrogen peroxide is made up of two hydrogen atoms and two atoms of oxygen. Similarly $C_6H_{12}O_6$ is molecular formula of glucose. It shows that every molecule of glucose has 6 carbon atoms, 12 hydrogen atoms and 6 oxygen atoms. The molecular formula of a compound shows the simplest ratio between different atoms present in the compound.

Benzene is a compound of carbon and hydrogen. It contains one C atom for every H atom. Each benzene molecule actually has six carbon atoms and six hydrogen atoms. Identify the empirical and molecular formulas of benzene from the following formulas.

$$C_6H_6 \quad , \quad CH$$

Molecular formulas for water and carbon dioxide are H_2O and CO_2 respectively. What are empirical formulas for these compounds?

For many compounds, empirical and molecular formulas are same. For example water (H_2O), carbon dioxide (CO_2) ammonia (NH_3), methane (CH_4), sulphur dioxide (SO_2) etc. Can you show it why?

Table 6.1 Formulae of some common elements and compounds

Element	Formula	Compound	Formula
Hydrogen	H_2	Water	H_2O
Oxygen	O_2	Carbon dioxide	CO_2
Nitrogen	N_2	Hydrochloric acid	HCl
Fluorine	F_2	Sodium hydroxide	$NaOH$
Chlorine	Cl_2	Copper (II) sulphate	$CuSO_4$
Bromine	Br_2	Glucose	$C_6H_{12}O_6$

CONCEPT ASSESSMENT EXERCISE 6.1

Write the empirical formulas for the compound containing carbon to hydrogen in the following ratios:

(a) 1:4 (b) 2:6 (c) 2:2 (d) 6:6

CONCEPT ASSESSMENT EXERCISE 6.2

1. Aspirin is used as a mild pain killer. There are nine carbon atoms, eight hydrogen atoms and four oxygen atoms, in this compound. Write its empirical and molecular formulas.
2. Vinegar is 5% acetic acid. It contains 2 carbon atoms, four hydrogen atoms and 2 oxygen atoms. Write its empirical and molecular formulas.
3. Caffeine ($C_8H_{10}N_4O_2$) is found in tea and coffee. Write the empirical formula for caffeine.

6.2 MOLECULAR MASS AND FORMULA MASS

Molecular mass is the sum of the atomic masses of all the atoms present in the molecule. All you have to do is to add up the atomic masses of all the atoms in the compound. For example,

Molecular mass of water H_2O = 2(atomic mass of H) + atomic mass of oxygen

= 2(1.008) + 16.00

= 2.016 + 16.00

= 18.016amu

Example 6.1: Determining molecular mass

1. Determine the molecular mass of glucose $C_6H_{12}O_6$ which is also known as blood sugar.
2. Determine the molecular mass of naphthalene $C_{10}H_8$, which is used in mothballs.

Problem solving strategy:

Multiply atomic masses of carbon, hydrogen and oxygen by their subscripts and add.

Solution:

1. Molecular mass of $C_6H_{12}O_6$ $= 6(12.00) + 12(1.008) + 6(16.00)$

   $= 180.096$ amu

2. Molecular mass of $C_{10}H_8$ $= 12 \times 10 + 1 \times 8$

   $= 120 + 8 = 128$ amu

The term molecular mass is used for molecular compounds. Whereas, the term formula mass is used for ionic compounds. Ionic compounds consist of arrays of oppositely charged ions rather than separate molecules. So we represent an ionic compound by its formula unit. A formula unit indicates the simplest ratio between cations and anions in an ionic compound. For example, the common salt consists of Na^+ and Cl^- ions. It has one Na^+ ion for every Cl^- ion. So formula unit for common salt is NaCl.

The sum of the atomic masses of all the atoms in the formula unit of a substance is called formula mass.

Example 6.2: Determining formula mass

1. Sodium Chloride, also called as table salt is used to flavour food, preserve meat, and in the preparation of large number of compounds. Determine its formula mass.
2. Milk of magnesia which contains $Mg(OH)_2$, is used to treat acidity. Determine its formula mass.

Problem solving strategy:

Add the atomic masses of all the atoms in the formula unit.

Solution:

1. Formula mass of NaCl $= 1 \times$ Atomic mass of Na $+ 1 \times$ Atomic mass of Cl

   $= 1 \times 23 + 1 \times 35.5$

   $= 58.5$ amu

2. Formula mass of $Mg(OH)_2$ $= 24 + 16 \times 2 + 1 \times 2$

   $= 24 + 32 + 2$

   $= 58$ amu

CONCEPT ASSESSMENT EXERCISE 6.3

1. Potassium Chlorate ($KClO_3$) is used commonly for the laboratory preparation of oxygen gas. Calculate its formula mass.
2. When baking soda, $NaHCO_3$ is heated carbon dioxide is released, which is responsible for the rising of cookies and bread. Determine the formula masses of baking soda and carbon dioxide.
3. Following compounds are used as fertilizers. Determine their formula masses.
   (i) Urea, $(NH_2)_2CO$ (ii) Ammonium nitrate, NH_4NO_3

6.3 CHEMICAL FORMULA AND NAME OF BINARY IONIC COMPOUNDS

A binary ionic compound is composed of mono-atomic metal cations and mono-atomic non-metal anions. To write the name of an ionic compound, the cation is named first followed by the name of anion. How do you name cations and anions? The name of the cation is the same as the name of the metal, but in the name of mono-atomic anion, the suffix ide is added to the root name of the element. For example, sodium chloride, magnesium oxide, aluminium nitride, etc.

The following steps are used to write the chemical formula of a binary ionic compound.

Step 1. Write the symbols for the cation first and then the symbols for the anion and their charges

Step 2. Balance the charges on the cations and anions using the smallest coefficient. The total charge on the cation must equal the total charge on the anion because an ionic compound is nutral

Step 3. Write coefficient as subscripts for each ion.

Step 4. Write the formula of ionic compound. For this leave out all the charges subscript which are 1.

For example, the chemical formula of sodium chloride is written as follows;

$$Na^{+1}Cl^{-1}$$

$$1(+1) = (-1)1$$

$$Na_1Cl_1 = NaCl$$

Example 2, the chemical formula of aluminium oxide is written as follows;

$$Al^{+3}O^{-2}$$

$$2(+3) = (-2)3$$

$$+6 = -6$$

$$Al_2O_3$$

Table indicating names and symbols for cations and anions

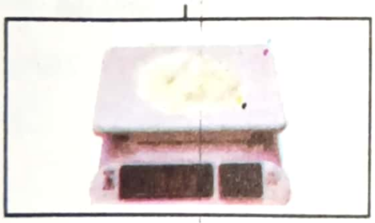
Cations		Anions	
Lithium	Li^+	Fluoride	F^-
Sodium	Na^+	Chloride	Cl^-
Ammonium	NH_4^+	Nitrite	NO_2^-
Potassium	K^+	Bromide	Br^-
Magnesium	Mg^{2+}	Nitrate	NO_3^-
Calcium	Ca^{2+}	Phosphate	PO_4^{3-}
Copper	Cu(II)	Sulfate	SO_4^{2-}

6.4 AVOGADRO'S NUMBER AND MOLE

How do you count your shoes? Since the shoes come in pairs, you'll probably count them in pairs rather than individually. Likewise, eggs, oranges, etc. are counted in the dozens, but paper with the ream. So the unit of counting depends on what you are counting. Chemists also use a practical unit for counting atoms, molecules and ions. They use a counting unit called mole to measure the amount of a substance.

A mole is an amount of a substance that contains 6.022×10^{23} particles of that substance. This experimentally determined number is known as Avogadro's number. It is represented by N_A. Just as a dozen eggs represent twelve eggs, a ream of paper represent 500 papers, a mole of a substance represents 6.022×10^{23} representative particles of a substance.. For example a mole of carbon is 6.022×10^{23} atoms. A mole of sulphur is 6.022×10^{23} atoms. A mole of water is 6.022×10^{23} molecules.

A mole of S-atoms
(6.022×10^{23} S-atoms)



A mole of C-atoms
(6.022×10^{23} C-atoms)

A pair of Shoes & a dozen eggs

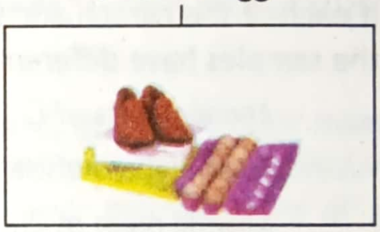


Figure 6.1 A mole of S-atoms, a mole of C-atoms & pair of Shoes & a dozen eggs

What is the mass of one mole C-atoms?

How many atoms are there in 32.1 g of S-atoms?

Does a dozen eggs have the same mass as a dozen bananas? Does a mole of carbon atoms have a different mass than a mole of sulphur atoms?

The mass of one mole of substance is called as molar mass. What are the molar masses of carbon and sulphur? The term representative particles in a substance are atoms, molecules, formula units or ions. For instance

Society, Technology and Science

Size of the Mole

Entire population cannot count 1 mole of coins in a year. They need about one million year to count them. So, when counting a pile of coins, it would not be convenient to count them one by one. The concept of mole has given a very simple method to count large number of items. Mole is not only a number but also represents definite amount of a substance. Just as 6.02×10^{23} carbon atoms weigh 12 g, 6.02×10^{23} coins will also have a definite mass. So an easy way is to weigh them. If you know the mass of one coin, you can count them by weighing.

water exists as molecules, therefore, one mole of water contains 6.022×10^{23} molecules of water. Hydrogen exists as H_2 molecules, so one mole of hydrogen contain 6.022×10^{23} molecules. Carbon exists as atoms so 1 mole of carbon contains 6.022×10^{23} atoms.

6.4.1 Gram Atomic Mass, Gram Molecular Mass and Gram Formula Mass

A mole of S-atoms
(6.022×10^{23} S-atoms)

What is the mass of 6.022×10^{23} S-atoms? Is this mass of S-atoms equal to its atomic mass?

A mole of C- atoms
(6.022×10^{23} C-atoms)

What is the mass of one mole of C-atoms? Is this mass of C-atoms equal to its atomic mass?

Atomic mass of an element expressed in grams is called gram atomic mass.

Is the gram atomic mass of C- atoms 12 g?

What is the gram atomic mass of S-atoms?

If each of the carbon and sulphur sample shown above contains one mole of atoms, why do the samples have different masses?

Atomic mass of C	= 12amu	\gram atomic mass of C = 12g
Atomic mass of Na	= 23amu	gram atomic mass of C = 23g
Atomic mass of Zn	= 63.54amu	gram atomic mass of C = 63.54g

Gram atomic mass of an element contains 1 mole of atoms.

Therefore,

Mass of 1 mole of C-atoms = 12g

Mass of 1 mole of Na-atoms = 23g

Mass of 1 mole of Zn-atoms = 63.54g

A mole of H_2O-molecules
(6.022×10^{23} H_2O-molecules)

A mole of $C_6H_{12}O_6$-molecules
(6.022×10^{23} $C_6H_{12}O_6$-molecules)

What is the mass of one mole of water molecules?
Is this mass of water molecules equal to molecular mass of water?

what is the mass of 6.022×10^{23} molecules of glucose?
is this mass of glucose molecules equal to molecular mass of glucose?

Molecular mass of a substance expressed in grams is called gram molecular mass.

	Molecular mass of H_2O	$= 2 \times 1.008 + 16$
		$= 18.016$amu
So,	gram molecular mass of H_2O	$= 18.016$g
	Molecular mass of $C_6H_{12}O_6$	$= 6 \times 12 + 12 \times 1.008 + 16 \times 6$
		$= 180.096$amu
So,	gram molecular mass of $C_6H_{12}O_6$	$= 180.096$g

Formula mass of a substance expressed in gram is called gram formula mass.

An ionic compound is represented by the formula unit that represents the simplest ratio between the ions of a compound. For example NaCl, KCl, $CuSO_4$ etc.

Formula mass of NaCl	$= 23 + 35.5$
	$= 58.5$amu
Therefore, gram formula mass of NaCl	$= 58.5$
Formula mass of KCl	$= 39 + 35.5$
	$= 74.5$amu
So, gram formula mass of KCl	$= 74.5$g

6.4.2 Difference between the Terms Gram Atomic Mass, Gram Molecular Mass And Gram Formula Mass

1. Gram atomic mass represents one mole of atoms of an element, gram molecular mass represents one mole of molecules of a compound or an element that exists in molecular state whereas gram formula mass represents one mole of ionic formula units of a compound.
2. Gram atomic mass contains 6.022×10^{23} atoms, gram molecular mass contains 6.022×10^{23} molecules whereas gram formula mass contain 6.022×10^{23} formula units.
3. All of these quantities represent molar mass. Mass of one mole of a substance expressed in grams is called molar mass. "Therefore, mole can be defined as atomic mass, molecular mass or formula mass expressed in grams".

6.5 CHEMICAL CALCULATIONS

In this section, you will learn about the chemical calculations based on the concept of mole and Avogadro's number.

6.5.1 Mole-Mass Calculations

Example 6.3: Calculating mass of one mole of a substance

Calculate the molar masses of (a) Na (b) Nitrogen (c) Sucrose $C_{12}H_{22}O_{11}$

Problem solving strategy:

If an element is a metal then its molar mass is its atomic mass expressed in grams (gram atomic mass). If an element exists as a molecule, its molar mass is its molecular mass expressed in grams (gram molecular mass).

Solution:

a) 1 mole of Na = 23g

b) Nitrogen occurs as a diatomic molecules.

Molecular mass of N_2 = 14 x 2

= 28amu

Therefore, mass of 1 mole of N_2 = 28 g

c) Molecular mass of $C_{12}H_{22}O_{11}$ = 12x12 + 1x22 + 16x11

= 144 + 22 + 176

Therefore, mass of 1mole of sucrose = 342g

CONCEPT ASSESSMENT EXERCISE 6.4

1. Calculate the mass of one mole of
   (a) Copper (b) Iodine (c) Potassium (d) Oxygen
2. Differentiate between gram formula mass and gram molecular mass.

Example 6.4: Calculating the mass of a given number of moles of a substance

Oxygen is converted to ozone (O_3) during thunder storms. Calculate the mass of ozone if 9.05 moles of ozone is formed in a storm?

Problem solving strategy:

Ozone is a molecular substance. Determine its molar mass and use it to convert moles to mass in grams.

9.05 moles of O_3 $\longrightarrow$? g of

Solution:

1 mole of O_3 = 16 x 3

= 48 g

1 mole of O_3 = 48 g

So, 9.05 moles of O_3 = 48 g x9.05

= 434.4g of O_3

Example 6.5: When natural gas burns CO_2 is formed. If 0.25 moles of CO_2 is formed, what mass of CO_2 is produced?

Problem solving strategy:

Carbon dioxide is a molecular substance. Determine its molar mass and use it to convert moles to mass in grams

0.25 moles of CO_2 ⟶ ? g of CO_2

Solution:

Molar mass of CO_2 = 12 + 16 x 2

= 44g

1 mole of CO_2 = 44g of CO_2

So, 0.25 moles of CO_2 = 44 x 0.25

= 11g of CO_2

Example 6.6: Converting grams to moles

How many moles of each of the following substance are present?

(a) A balloon filled with 5g of hydrogen.

(b) A block of ice that weighs 100g.

Problem solving strategy:

Hydrogen and ice both are molecular substances. Determine their molar masses. Use the molar mass of each to convert the masses in grams to moles.

mass ⟶ ? moles

Solution:

a) Molar mass of H_2 = 1.008 x 2

= 2.016g

1 mole of H_2 = 2.016g

So, 2.016g of H_2 = 1 mole of H_2

1g of H_2 = $\frac{1}{2.016}$ moles of H_2

5g of H_2 = $\frac{1}{2.016}$ x5 moles of H_2

= 2.48 moles of H_2O

b) 1 mole of H_2O = 2 x 1.008 + 16

= 2.016 + 16

= 18.016g

1 mole of H_2O = 18.016g

So, 18.016g of H_2O = 1 mole

1g of H_2O = $\frac{1}{18.016}$ moles

100g of H_2O $= \frac{1}{18.016} \times 100$ moles
$= 5.55$ moles of H_2O

CONCEPT ASSESSMENT EXERCISE 6.5

1. The molecular formula of a compound used for bleaching hair is H_2O_2. Calculate (a) Mass of this compound that would contain 2.5 moles. (b) No. of moles of this compound that would exactly weigh 30g.
2. A spoon of table salt, NaCl contains 12.5grams of this salt. Calculate the number of moles it contains.
3. Before the digestive systems X-rayed, people are required to swallow suspensions of barium sulphate $BaSO_4$ Calculate mass of one mole of $BaSO_4$.

6.5.2 Mole-Particles Calculations

Example 6.7: Calculating the number of atoms in given moles

1. Zn is a silvery metal which is used to galvanize steel to prevent corrosion. How many atoms are there in 1.25 moles of Zn.
2. A thin foil of aluminium (Al) is used as wrapper in food industries. How many atoms are present in a foil that contains 0.2 moles of aluminium?

Problem solving strategy:

Remember that symbols Zn and Al stand for one mole of Zn and Al atoms respectively.

Solution:

1. 1 mole of Zn contains $= 6.022 \times 10^{23}$ atoms
   1.25 moles of Zn contain $= 6.022 \times 10^{23} \times 1.25$
   $= 7.53 \times 10^{23}$ Zn atoms
2. 1 mole of Al contains $= 6.022 \times 10^{23}$ atoms
   So 0.2 moles of Al will contain $= 6.022 \times 10^{23} \times 0.2$
   $= 1.2044 \times 10^{23}$ atoms

Example 6.8: Calculating the number of molecules in given moles of a substance

1. Methane (CH_4) is the major component of natural gas. How many molecules are present in 0.5 moles of a pure sample of methane?
2. At high temperature hydrogen sulphide (H_2S) gas given off by a volcano is oxidized by air to sulphur dioxide (SO_2). Sulphur dioxide reacts with water to form acid rain. How many molecules are there in 0.25 moles of SO_2

Problem solving strategy:

Remember that CH_4 is a molecular compound, thus 1 mole of methane will have 6.022×10^{23}

molecules. Similarly, SO_2 is a molecular compound, its one mole will also have 6.022×10^{23} molecules.

Solution:

1. 1 mole of CH_4 contains $= 6.022 \times 10^{23}$ molecules

So, 0.5 moles of CH_4 will contain $= 6.022 \times 10^{23} \times 0.5$

$= 3.011 \times 10^{23}$ molecules

2. 1 mole of SO_2 contains $= 6.022 \times 10^{23}$ molecules

So, 0.25 moles of SO_2 will contain $= 6.022 \times 10^{23} \times 0.25$

$= 1.5055 \times 10^{23}$ molecules

Example 6.9: Calculating the number of moles in the given number of atoms

Titanium is corrosion resistant metal that is used in rockets, aircrafts and jet engines. Calculate the number of moles of this metal in a sample containing 3.011×10^{23} Ti-atoms.

Problem solving strategy:

Remember that 1 mole of an element contains 6.022×10^{23} atoms.

Thus,

6.022×10^{23} atoms = 1 mole

3.011×10^{23} atoms $\longrightarrow$? moles

Solution:

6.022×10^{23} Ti atoms = 1 mole of Ti

1 Ti atom $= \frac{1}{6.022 \times 10^{23}}$ moles of Ti

3.011 x 1023 Ti atoms $= \frac{1}{6.022 \times 10^{23}} \times 3.011 \times 10^{23}$ moles of Ti

= 0.5 moles of Ti

Example 6.10: Calculating number of moles in the given number of molecules

Formaldehyde is used to preserve dead animals. Its molecular formula is CH_2O Calculate the number of moles that would contain 3.011×10^{22} molecules of this compound.

Problem Solving Strategy:

Remember that 1 mole of any compound contains 6.022×10^{23} molecules.

Thus,

6.022×10^{23} molecules = 1 mole of compound

3.011×10^{22} molecules $\longrightarrow$? moles

Solution:

6.022×10^{23} molecules = 1 mole of formaldehyde

1 molecule $= \frac{1}{6.022 \times 10^{23}}$ moles of formaldehyde

3.011×10^{22} molecules $= \frac{1}{6.022 \times 10^{23}} \times 3.011 \times 10^{22}$ moles of formaldehyde
$= 0.05$ moles of formaldehyde

CONCEPT ASSESSMENT EXERCISE 6.5

1. Aspirin is a compound that contains carbon, hydrogen and oxygen. It is used as a painkiller. An aspirin tablet contains 1.25×10^{20} molecules. How many moles of this compound are present in the tablet?
2. A method used to prevent rusting in ships and underground pipelines involves connecting the iron to a block of a more active metal such as magnesium. This method is called cathodic protection. How many moles of magnesium are present in 1 billion (1×10^{9}) atoms of magnesium.

6.6 CHEMICAL EQUATION AND BALANCING

The symbolic representation of a chemical reaction is called chemical equation. The reactants in a chemical equation are the substances that initiate the chemical reaction, and the products are the substances that are formed during the chemical reaction. Reactants are always written on the left of the equation and products on the right, an arrow between them is used to show the direction of the chemical change.

For writing a chemical equation, follow the following steps.

Step 1. Identify reactants and products and write word equation for the reaction. Represent chemical equation is as follows:

Reactants → Products

Step 2. Write the symbols and formulae of reactants and products. Indicate their physical states in parenthesis. Use s for solid, l for liquid, g for a gas and aq for aqueous.

Example 6.11: burning of coal is represented as follows.

Coal + Oxygen → Carbon dioxide

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

CONCEPT ASSESSMENT EXERCISE 6.6

Represent the following chemical reactions by chemical equations.

1. Burning of hydrogen(H_2) to produce water.
2. Burning of magnesium(Mg) to produce magnesium oxide(MgO).

6.6.1 Balancing a chemical equation

A chemical reaction only changes the arrangements of atons. The number of atons remains the same. Count the number of atoms of atoms of each type in the following equation:

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

Reactants	Products
1 C-atom	1 C-atom

2 O-atoms 2 O-atoms

Note that the number of atoms of each type are the same on the reactant side adn the product side. Such a chemical equation is called a balanced chemical equation.How can you balance a chemical reaction, it is unbalanced?

Consider the following reaction;

$$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

Reactants	Products
1 C-atom	1 C- atom
4 H-atoms	2 H-atoms
2 O-atoms	3 O-atoms

C-atoms are balanced, but H and O- atoms are unbalanced.

Balance one element at a time. To balance the chemical equation use co-efficients. Always start with the lowest co-efficient. Remember that you should not chamge **subscripts in a chemical formula.**

Step 1

Put a co-efficient 2 before H_2O to balance H- atoms.

$$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

Step 2

Now balance O-atoms. There are 2O-atoms on the left side and O-atoms on the right side. Put 2 before O_2

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

Step 3

Now check the equation again, it is balanced.

6.6.2 Exploring ionic equation

A chemical equation in which substances dissolved in water are written as individual ions is called an ionic equation. For example,

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

Write the substances that are soluble in water in their dissociated form.

$$H^{+1}(aq) + Cl^{-1}(aq) + Na^{+1}(aq) + OH^{-1}(aq) \rightarrow Na^{+1}(aq) + Cl^{-1}(aq) + H_2O(l)$$

Remove common ions from both sides. These ions do not actually take part in the chemical reaction and are called spectator ions. Write net ionic equation.

$$H^{+1}(aq) + OH^{-1}(aq) \rightarrow H_2O(l)$$

Example 6.12: Transform the following chemical equations into ionic equations.

$$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_4 + H_{2(g)}$$

$$Mg_{(s)} + 2H^{1+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2H^{1+}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + H_{2(g)}$$

CONCEPT ASSESSMENT EXERCISE 6.7

Transform the following chemical equations into ionic equations.

1. $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$
2. $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

6.7 MOLECULAR AND STRUCTURAL FORMULA

A structural formula of a compound shows the arrangement of atoms present in it. Whereas a molecular formula shows the number of atoms of each element. For example,

Structural formula of n-Butane is CH_3-CH_2-CH_2-CH_3 and its molecular formula is C_4H_{10}.

How can you write the molecular formula of a compound from its structural formula?

Follow the following steps:

1. Identify different types of elements present in the structural formula.
2. Write symbols of these elements side by side in a line.
3. Count the number of atoms of each element from the structural formula.
4. Show this number of atoms as subscripts of symbol of corrosponding element.

Example 6.13: Writing the molecular formula of the following compound.

CH_3-CH_2-CH_2- OH

Solution: C_3H_8O

CONCEPT ASSESSMENT EXERCISE 6.8

Write the molicular formulae of the following compounds,

1. CH_3-CH_2-OH
2. CH_3-CH-$_2$-NH_2
3. CH_3-CO-CH_3

KEY POINTS

- Chemistry is the science of materials of the universe.
- The branch of Chemistry that deals with laws and theories to understand the structure and changes of matter is called Physical Chemistry.
- An element is a substance all the atoms of which have the same atoı ic number.
- A compound consists of two or more elements held together in fix· ɹroportions by chemical bonds.
- Chemical formula of a compound that gives the simplest whole-number ratio between atoms is called empirical formula.

- Molecular formula of a compound gives the exact number of atoms present in a molecule.
- Molecular mass is the sum of atomic masses of all the atoms present in the molecule.
- The number of representative particles in one mole of the substance is known as Avogadro's number.
- The amount of matter that contains as many atoms, ions or molecules as the number of atoms in exactly 12g of C-12 is called mole. Mole can also be defined as atomic mass, molecular mass or formula mass expressed in grams.
- Atomic mass of an element expressed in grams is called gram atomic mass.
- Molecular mass of an element or a compound expressed in grams is its gram molecular mass.
- Gram formula mass is the formula mass of a substance in grams.

References for additional information

- Zumdahl, Introductory Chemistry.
- Raymond Chang, Essential Chemistry.

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) What is the formula mass of $CuSO_4.5H_2O$. (Atomic masses: Cu=63.5, S=32, O=16, H=1)

(a) 159.5 (b) 185.5
© 249.5 (d) 149.5

(ii) A compound with chemical formula Na_2CX_3 has formula mass 106amu. Atomic mass of the element X is

(a) 106 (b) 23
(c) 12 (d) 16

(iii) How many moles of molecules are there in 16g oxygen.

(a) 1 (b) 0.5
© 0.1 (d) 0.05

(iv) What is the mass of 4 moles of hydrogen gas.

(a) 8.064g (b) 4.032g
(c) 1g (d) 1.008g

(v) What is the mass of carbon present in 44g of carbon dioxide.

(a) 12g (b) 6g
(c) 24g (d) 44g

(vi) Which term is the same for one mole of oxygen and one mole of water?

(a) volume (b)
(c) mass (d) atoms
(e) molecules

(vii) If one mole of carbon contains x atoms, what is the number of atoms contained in 12g of Mg.

(a) x (b) 0.5x

(c) 2x (d) 1.5x

2. Give short answer.

(i) What is mole?

(ii) Differentiate between empirical formula and molecular formula.

(iii) What is the number of molecules in 9.0 g of steam?

(iv) What are the molar masses of uranium -238 and uranium -235?

(v) Why one mole of hydrogen molecules and one mole of H-atoms have different masses?

3. Define ion, molecular ion, formula unit, free radical, atomic number, mass number, atomic mass unit.

4. Describe how Avogadro's number is related to a mole of any substance.

5. Calculate the number of moles of each substance in samples with the following masses:

(a) 2.4 g of He (b) 250mg of carbon

(c) 15g of sodium chloride (d) 40g of sulphur

(e) 1.5kg of MgO

6. Calculate the mass in grams of each of the following samples:

(a) 1.2 moles of K (b) 75moles of H_2

(c)0.25 moles of steam (d) 1.05 moles of $CuSO_4.5H_2O$

(e) 0.15moles of H_2SO_4

7. Calculate the number of molecules present in each of the following samples:

(a) 2.5 moles of carbon dioxide (b) 3.4 moles of ammonia, Nh_3

© 1.09 moles of benzene, C_6H_6 (d)0.01 moles of acetic acid, CH_3COOH

8. Decide whether or not each of the following is an example of empirical formula:

(a) Al_2Cl_6 (b) Hg_2Cl_2

(c)NaCl (d) C_2H_6O

9. TNT or trinitrotoluene is an explosive compound used in bombs. It contains 7 C-atoms, 5 H-atoms, 3-Natoms and 6 O-atoms. Write its em irical formula.

10. A molecule contains four phosphorus atoms and ten oxygen atoms. Write the empirical formula of this compound. Also determine the molar mass of this molecule.

11. Indigo ($C_{16}H_{10}N_2O_2$), the dye used to colour blue jeans is derived from a compound

known as indoxyl (C_8H_7ON). Calculate the molar masses of these compounds. Also write their empirical formulas.

12. Identify the substance that has formula mass of 133.5amu.

(a) $MgCl_2$ (b) S_2Cl_2

(c) BCl_3 (d) $AlCl_3$

13. Calculate the number of atoms in each of the following samples:

(a) 3.4 moles of nitrogen atoms (b) 23g of Na

(c) 5g of H atoms

14. Calculate the mass of the following:

(a) 3.24×10^{18} atoms of iron (b) 2×10^{10} molecules of nitrogen gas

(c) 1×10^{25} molecules of water (d) 3×10^{6} atoms of Al

15. Balance the following chemical equations

(a) $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$

(b) $NH_{3(g)} \rightarrow N_{2(g)} + H_{2(g)}$

16. Potassium is Group 1 element. It is silvery white metal. It burns in air and forms both potassium oxide and potassium nitride. The nitride ion is N^{3-}.

(a) Predict the formula of potassium oxide and potassium nitride.

(b) A 0.5g sample of K was added in 100cm^3 of water.

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

Show that 28x10^{-2} mole of K were added to the water.

(c) Balance above chemical equation.

(d) Transform above chemical equation into ionic equation.

(e) Calculate the number of atoms present in the sample of K.

(f) Predict period number of potassium in the periodic table.

UNIT 07

ELECTROCHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define redox reactions as simultaneous oxidation and reduction in terms of transfer of oxygen, hydrogen, electrons, and changes in oxidation state.
- Use roman numerals to indicate oxidation number of an element in a compound.
- Identify oxidizing and reducing agents in redox reactions.
- Recognize that the oxidation number of elements in their free state is zero.
- Derive the formula of ionic compounds from ionic charges and oxidation numbers.
- Identify the oxidation number of a monoatomic ion is the same as the charge on the ion.
- Explain that the sum of the oxidation numbers in a neutral compound is zero.
- Explain that the sum of the oxidation numbers in an ion is equal to the charge on the ion.
- Identify the redox reaction by the colour changes involved when using acidified aqueous potassium manganate(VII) to (II) or aqueous potassium iodide.
- Define corrosion and discuss methods to prevent it. (Some examples may include barrier method such using paint, galvanizing, electroplating, sacrificial protection such as using magnesium blocks in ships.

INTRODUCTION

Oxidation-reduction reactions (redox) are fundamental chemical processes that play a crucial role in a variety of natural phenomena and industrial applications. This chapter examines the commonalities between different redox reactions, such as the rusting of iron objects, the burning of fuel in car engines, forest fires, and the metabolism of food in human and animal bodies. In addition, the importance of redox reactions in the production of electricity in batteries, in the decolorization of substances with household bleaches, and in the industrial production of important metals and chemicals is discussed.

One of the most important application of electrochemistry is batteries and fuels cells, which use chemical energy to generate electrical energy.

7.1 OXIDATION AND REDUCTION

We can define redox reactions in terms of transfer of oxygen, hydrogen, and electrons. In redox reactions oxidation and reduction occur simultaneously.

7.1.1 Oxidation-Reduction in Terms of Loss or Gain of Oxygen, Hydrogen or Electrons

Oxidation is gain of oxygen and reduction is loss of oxygen

In steel mills iron ores, usually oxides of iron are converted to the pure metal commercially by the reaction with coke (carbon) in the blast furnace. The carbon first reacts with air to form carbon monoxide, which in turn reacts with iron oxide as follows.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Which substance . this reaction is losing oxygen? Which substance is gaining oxygen?

CO is gaining oxygen, so undergoing oxidation.

Fe_2O_3 is losing oxygen, so undergoing reduction.

7.1.2 Oxidation-reduction in terms of transfer of hydrogen

Oxidation is loss of hydrogen and reduction is gain of hydrogen

Acetylene (C_2H_2) is commercially used for cutting and welding metals. When acetylene burns, it produces a very hot flame known as oxy-acetylene flame. Following reaction takes place when it burns.

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

Which substance is losing hydrogen? Which substance is gaining hydrogen?

C_2H_4 is losing hydrogen, so unregoing oxidation.

O_2 is gaining hydrogen, so undergoing reduction.

7.1.3 Oxidation and Reduction in Terms of Transfer of Electrons

Oxidation is loss of electrons and reduction is gain of electrons.

For example, consider the following reaction.

$$CuO + Zn \longrightarrow Cu + ZnO$$

Copper(II)oxide and zinc(II)oxides are both ionic compounds. Rewrite this equation as ionic equation.

$$Cu^{2+} + O^{2-} + Zn \longrightarrow Cu + Zn^{2+} + O^{2-}$$

Oxide ions are spectator ions, so net equation is;

$$Cu^{2+} + Zn \longrightarrow Cu + Zn^{2+}$$

Cu^{2+} ions are gaining 2 electrons to form Cu. Its oxidation state changes from +2 to zero. So Cu^{2+} are undergoing reduction. The oxidation number of Zn is increasing from zero to +2, so Zn is losing 2 electrons and undergoing oxidation.

- Oxidation is defined as the loss of hydrogen, gain of oxygen or loss of electrons.
- Reduction is defined as the gain of hydrogen, loss of oxygen or gain of electrons.

Example 7.1: Identifying the element undergoing oxidation in terms of transfer of oxygen or hydrogen

Following reaction occurs when you burn Sui gas.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + heat$$

Identify the element undergoing oxidation.

Problem solving strategy:

Identify the substance that gains O-atoms or loses H-atoms.

Solution:

Since C in CH_4 loses H-atoms and combines with oxygen atoms, thus C atoms undergo oxidation. At the same time O-atoms combine with H-atoms to form H_2O, thus O-atoms undergo reduction

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Redox in photography

A photographic film is basically an emulsion of silver bromide, (AgBr) in gelatin. When the film is exposed to light, Silver bromide granules become activated. This activation depends on the intensity of the light falling upon them. When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is a mild reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver. Reduced silver atoms form image.

$$Ag^+ + 1e^- \longrightarrow Ag_{(s)}$$

Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to the light appear darkest because they have the highest concentration of metallic silver. Thus photography involves oxidation-reduction reaction.

CONCEPT ASSESSMENT EXERCISE 7.1

Identify elements undergoing oxidation and reduction in the following reactions:

1. $N_2 + 3H_2 \longrightarrow 2NH_3$
2. $2H_2 + O_2 \longrightarrow 2H_2O$
3. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
4. $4Al + 3O_2 \longrightarrow 2Al_2O_3$

Example 7.2: Identifying the element oxidized or reduced in terms of transfer of electrons

In the following reaction identify which element is oxidized and which element is reduced

$$2Ca + O_2 \longrightarrow 2CaO$$

Problem solving strategy:

Ca being metal forms cation by losing electrons (oxidation) and oxygen being non-metal gains electrons (reduction) to form anion.

Solution:

Remember that Group IIA metals form M^{2+} cations, and that Group VIA non-metals form X^{2-} anions. This means in this reaction each Ca atom loses two electrons to form Ca^{2+}, so it is oxidized. Each oxygen atom gains two electrons to form O^{2-}, so it is reduced.

$$2Ca \longrightarrow 2Ca^{2+} + 2e^- \quad \text{(oxidation)}$$

$$O_2 + 4e^- \longrightarrow 2O^{2-} \quad \text{(reduction)}$$

CONCEPT ASSESSMENT EXERCISE 7.2

In the following reactions, identify which element is oxidized and which element is reduced in terms of electron transfer.

$$4Na + O_2 \longrightarrow 2Na_2O$$

$$2Al + 3Cl_2 \longrightarrow 2AlCl_3$$

$$Mg + Cl_2 \longrightarrow MgCl_2$$

7.2 OXIDATION STATES AND RULES FOR ASSIGNING OXIDATION STATES

7.2.1 Oxidation States

Oxidation state or oxidation number is defined as the number of charges an atom will have in a molecule or a compound.

The elements that show an increase in oxidation number are oxidized. The elements that show a decrease in oxidation number are reduced. Do you think H in HCl is oxidized and Cl is reduced? Comparison of oxidation and reduction processes is given in table 7.1.

Table 7.1: Process leading to oxidation and reduction	
Oxidation	**Reduction**
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
Increase in oxidation number	Decrease in oxidation number

7.2.2 Rules for Assigning Oxidation States or Numbers

1. The oxidation state of any uncombined or free elements is always zero e.g., oxidation state of Zn, Na, H in H_2, S in S_8 etc is zero.
2. In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^{1+} and Ca in Ca^{2+} is +1 and +2 respectively.
3. In a complex ion the sum of oxidation states of atoms is equal to the charge on their ion. e.g., in CO_3^{2-}, the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^{1+}, the sum of oxidation states of N and 4H atoms is +1.
4. The oxidation number of each of the atoms in a molecule or compound is counted separately and their algebraic sum is zero e.g., In HCl, the sum of oxidation states of H and Cl atoms is zero. Similarly in CO_2, the sum of oxidation states of one C and 2 oxygen atoms is zero.

Table 7.2 shows the oxidation states of some of the elements in binary compounds which rarely change.

Table 7.2: Oxidation states of some elements in binary compounds that rarely change

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2(except peroxides and in OF_2)

Monoatomic ions and their oxidation numders

The oxidation number of a monatomic ion is equal to its charge. For example, Na^+ is formed after a Na atom has lost one electron to gain a +1 charge. So its oxidation number is +1. Similarly, a chlorine atom forms a Cl^- ion after gaining one electron to obtain a -1 charge. So its oxidation number is -1. The oxidation number of an atom is the number of electrons the atom has lost or gained. Because a monatomic ion is formed by the gain or loss of electrons from a single atom, its charge is equal to its oxidation number..

Polyatomic ions and their oxidation numders

In a polyatomic ion, the sum of the oxidation numbers of all the atoms is equal to the charge on the ion. For example, in CO_3^{2-} ion the oxidation numbers of carbon and oxygen are +4 and -2 respectively. So, the sum of the oxidation number of one carbon atom and three oxygen atoms would be $1(+4) + 3(-2) = -2$ which is the charge on the ion.

7.2.3 Determining the Oxidation Number of an Atom in a Compound

Let's see how to use rules discussed in section 7.2.2 to determine the oxidation number of an atom of an element in a compound.

Example 7.3: Determining oxidation number

A device called Breathalyzer is used by police to test a person's breath for alcohol. It contains an acidic solution of potassium dichromate $K_2Cr_2O_7$. It is a strong oxidizing agent. Determine oxidation state of Cr in it.

Problem Solving Strategy:

Use rules 1 to 4 and table 7.1 to get as many oxidation numbers as you can. Use rule 4 to get oxidation number that has not been assigned.

Solution:

1. The oxidation number of K is +1, since it belongs to Group-1A. There are 2 K atoms therefore, overall oxidation number for K is 2(+1) =+ 2
2. There are 7 oxygen atoms, therefore overall oxidation state for O is 7(-2) = -14
3. Suppose oxidation for Cr is x, since there are two Cr atoms, therefore, overall oxidation sate for Cr is 2x.
4. The sum of oxidation numbers must be zero.

$$+2 + 2x + (-14) = 0$$
$$2x - 12 = 0$$
$$2x = 12$$
$$x = +6$$

Thus oxidation state for Cr in $K_2Cr_2O_7$ is +6

Example 7.4: Determining oxidation state

Boric acid H_3BO_3 is used in eye wash. What is the oxidation state of B in this acid?

Problem solving strategy:

Use rules and table 7.2 to get the oxidation state of H and O- atoms. Use rule 4 to get the oxidation state of B.

Solution:

1. There are 3 H-atoms, therefore, overall oxidation state for H is

$$3(+1) = +3$$

2. There are 3 O-atoms, therefore, overall oxidation state for O is

$3(-2) = -6$

3. Suppose the oxidation state for B is x.

4. The total oxidation states for all the atoms must be zero.

$+3 + x + (-6) = 0$

$+3 + x - 6 = 0$

$x - 3 = 0$

$x = 3$

Thus the oxidation state for B in H_3BO_3 is + 3.

CONCEPT ASSESSMENT EXERCISE 3.1

One major problem of air pollution is the formation of acid rain. Air pollutants such as SO_2 and NO_2 combine with oxygen and water vapours in the air to form H_2SO_4 and HNO_3. These acids fall to the ground with the rain, making the rain acidic. Clouds can also absorb the acids and carry them hundreds of kilometers away from where the pollutants are released. Determine the oxidation number of N in NO_2 and HNO_3, S in SO_2 and H_2SO_4

Example 7.5: Determining the oxidation number of an element in an ion.

What is the oxidation number of C in carbonate ion, CO_3^{-2}

Problem Solving Strategy:

(a) Use rule that oxidation number of O is -2

(b) Use rule 3 to find oxidation state of C

Solution:

Oxidation State of one C-atom + Oxidation state of 3-O atoms = -2

Oxidation state of C-atom + 3(-2) = -2

Oxidation state of C-atom -6 = -2

Oxidation state of C-atom =-2+6

Oxidation state of C-atom = +4

Thus the oxidation of C in carbonate ion is +4

CONCEPT ASSESSMENT EXERCISE 7.4

Determine the oxidation state of

1. S in sulphate ion, SO_4^{-2}
2. P in phosphate ion, PO_4^{-3}
3. N in ammonium ion, NH_4^{+1}

7.3 FORMULA OF AN IONIC COMPOUND

To determine the formula of an ionic compound from the ionic charges and oxidation numbers of the constituent ions, you must determine the simplest whole-number ratio of cations (positively charged ions) to anions (negatively charged ions) that results in a neutral compound. Ionic compounds are electrically neutral, meaning that the total positive charge of the cations is equal to the total negative charge of the anions. Let's go through step by step to derive the formula:

Example: Calcium chloride

Step 1: Identify the ions in the compound and their charges. Consider, for example, the combination of calcium ions (Ca^{2+}) and chloride ions (Cl^{-}).

Step 2: Determine the ratio of charges needed to balance each other. In this case, calcium has a charge of +2 and chloride has a charge of -1. You need two chloride ions for every calcium ion to balance the charges.

Step 3: Write a formula with assignments that represents the relationship defined in Step 2. The formula for calcium chloride is $CaCl_2$.

Example: Magnesium oxide

Step 1: Identify the ions and their charges - magnesium ions (Mg^{2+}) and oxide ions (O^{2-})

Step 2: Determine the charge - the magnesium has a charge of 2 and the oxide has a charge of -2. Thus, one magnesium ion is required for each oxide ion.

Step 3: Write the formula - MgO.

Example: Aluminum sulphate

Step 1: Identify the ions and their charges—aluminum ions (Al^{3+}) and sulphate ions (SO_4^{2-})

Step 2: Determine the charge - aluminum has a charge of +3 and sulphate has a charge of -2. You need two aluminum ions for every three sulphate ions to balance the charges.

Step 3: Write the formula - $Al_2(SO_4)_3$.

It is important to remember that when writing a formula, parentheses must be used to add polyatomic ions when more than one is needed to balance the charges. By following these steps and understanding the charges on the ions, you can derive the formulas of various ionic compounds.

7.3.1 Use of Roman numerals as oxidation number

Roman numerals are used to indicate the oxidation states of elements in compounds, when a metal exhibits variable oxidations in compounds. For examples in transition metal compounds.

$CuSO_4$ is written as copper (II) sulphate.

$FeCl_2$ is written as iron (II) chloride.

$FeCl_3$ is written as iron (III) chloride

7.4 OXIDIZING AND REDUCING AGENTS

An oxidizing agent is a substance that causes another substance to oxidize by taking electrons from it. It is often called an electron acceptor because it accepts electrons during a reaction.

Oxidizing agents themselves are reduced in the process (they gain electrons). Examples of common oxidizing agents are oxygen (O_2), hydrogen peroxide (H_2O_2), chlorine (Cl_2) and potassium permanganate ($KMnO_4$). **A reducing agent is a substance that causes another substance to be reduced by donating electrons to it.** It is often called an electron donor because it loses electrons during the reaction. The reducing agents themselves are oxidized in the process (they lose electrons). Examples of common reducing agents include hydrogen gas (H_2), metal hydrides (such as $NaBH_4$), carbon monoxide (CO), and metals such as zinc (Zn) and aluminum (Al). it.

For example, in the reaction between sodium and chlorine to form sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Na is reducing agent as it is oxidize vhereas Cl_2 is oxidizing agent as it is reduced.

Activity 7.1

Prepare solutions of ferrous sluphate($FeSO_4$) and potassium permanganate ($KMnO_4$) in separate beakers. Transfer about 10cm^3 of ferrous sulphate solution in a test tube. Add about 10cm^3 of dill. H_2SO_4 in it. Now add few drops of $KMnO_4$solution in the test tube. What happens?

$FeSO_4$reduces $KMnO_4$, so its purple colour is discharged. $KMnO_4$ oxidizes $FeSO_4$ in this reaction.

$FeSO_4$ is reducing agent whereas $KMnO_4$is oxidizing agent

A color change during a chemical reaction may indicate a redox reaction.

Potassium permanganate (VII) is an oxidizing agent.

Often used to test for the presence of reducing agents.

When acidified $KMnO_4$ is added to the reducing agent, it changes from purple to colorless.

The above reaction discharges the purple colour of $KMnO_4$

Therefore, the solution must contain a reducing agent that reduces MnO_4^{2-} ions (purple) to Mn^{2+} ions (colourless).

During this reaction, the oxidation number of Mn changes from +7 to +2.

Activity 7.2

Prepare an aqueous solution of potassium iodide and transfer it to a 10 cm^3 test tube. Add about 5 cm^3 of hydrogen peroxide to it. what's going on, the solution turns reddish-brown, indicating the formation of iodine.

Potassium iodide is a reducing agent. Often used to test for the presence of oxidizing agents. When a potassium iodide solution is added to an acidified hydrogen peroxide solution, the solution turns reddish-brown. The appearance of this color is due to the formation of iodine I_2.

KI is oxidized and H_2O_2 is reduced.

This redox reaction can be confirmed by a color change from colorless to reddish-brown.

$$H_2O_{2(l)} + 2KI_{(aq)} + 2H^+_{(aq)} \longrightarrow 2K^+_{(aq)} + I_{2(aq)} + 2H_2O$$

7.4.2 How can you identify oxidizing and reducing agents in a chemical reaction?

Consider the following reaction that takes place in the manufacture of steel.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

To identify the oxidizing and reducing agents, work out the oxidation states of all the elements involved in the reaction.

$$\overset{2(+3)\ (-2)3}{Fe_2O_3} + \overset{+2-2}{3CO} \longrightarrow \overset{0}{2Fe} + \overset{+4(-2)2}{3CO_2}$$

(i) Carbon is being oxidized because there is an increase in its oxidation state.

(ii) Fe is being reduced because there is a decrease in its oxidation state.

(iii) The reactant CO contains the C that is being oxidized, so CO is reducing agent.

(iv) The reactant Fe_2O_3 contains the Fe that is being reduced. So Fe_2O_3 is oxidizing agent.

Oxidizing or reducing agent is the whole molecule or formula unit and not the atom that has undergone change in oxidation number.

Example 7.6: Identifying the oxidizing and reducing agents

Tungsten is used to make filaments for electric bulbs because it has the highest melting point and high electrical resistance. This metal is obtained from tungsten (VI) oxide, WO_3 by reducing it with hydrogen gas.

$$WO_3 + 3H_2 \longrightarrow W + 3H_2O$$

Identify the oxidizing and reducing agents in this reaction.

Problem solving strategy:

Step 1: Workout the oxidation states of all the elements involved in the reaction.

Step 2: Note the element that is undergoing an increases in its oxidation state. Since it is being oxidized. The reactant that contains this element is reducing agent.

Step 3: Note the element that is undergoing a decrease in its oxidation state. Since it is being reduced. The reactant that contains this element is oxidizing agent.

Solution:

First assign oxidation numbers to each atom.

$$\overset{+6\ (-2)3}{WO_3} + \overset{0}{3H_2} \longrightarrow \overset{0}{W} + \overset{2(+1)\ -2}{3H_2O}$$

Because the oxidation number of W decreases, so WO_3 is an oxidizing agent. Similarly the oxidation number of H increases, therefore H_2 is reducing agent.

CONCEPT ASSESSMENT EXERCISE 7.5

1. **Identify oxidizing and reducing agents in the following reactions:**
   a) $2S + Cl_2 \longrightarrow S_2Cl_2$
   b) $2Na + Br_2 \longrightarrow 2NaBr$
2. **Differentiate between oxidizing and reducing agents.**

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Redox in photography

Silver is very soft metal. Silver atoms have weak interactions and are loosely packed together. Silver tarnishes in air when it comes in contact with trace quantities of H_2S or SO_2 in the atmosphere or food such as eggs, that are rich in sulphur compounds. Silver tarnish is silver sulphide that gives silver blackish appearance. Due to this reason decorative and practical objects made of solid silver gradually turn black and lose shining appearance. Decorative and practical objects are plated with a thin layer of silver. Atoms in thin layers firmly adhere to the metal atoms of the object and form a durable layer. An article thickly plated with silver contains many layers of silver atoms. Such layers form soft covering. These layers gradually turn black.

7.5 CORROSION AND ITS PREVENTION

7.5.1 Corrosion

Corrosion is a natural electrochemical process that occurs when a metal reacts with its environment. **In this reaction a metal reacts with oxygen and moisture in the atmosphere.** Corrosion converts refined metals to the more stable metal oxides. It can cause significant damage to structures, vehicles, and equipments.

Most familiar example of corrosion is the formation of rust on iron. Oxygen and water are necessary for iron to rust. Corrosion is an oxidation-reduction reaction. A region of metal surface that has relatively less moisture, acts as anode. Will Fe oxidize in this region?

$$Fe\,(s) \longrightarrow Fe^{+2}(aq) + 2e^-$$

Another region on the surface of metal that has relatively more moisture acts as cathode. The electrons released in the oxidation process reduce atmospheric oxygen to hydroxyl ions.

$$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$$

The Fe^{+2} ions formed at the anodic regions flow to the cathodic regions through the moisture on the surface. Here Fe^{+2} ions further react with oxygen to form rust, $Fe_2O_3.xH_2O$

7.5.2 Prevention of Corrosion

Corrosion is a widespread issue that affects industries, infrastructures, and everyday objects. Therefore, understanding corrosion and implementing preventive measures are crucial. Prevention of corrosion is an important way of conserving our natural resources. Following methods have been devised to protect metals from corrosion:

1. **Coating with paint:**

   Corrosion can be prevented by applying protective coating such as paint or epoxy creating a barrier between metal surface and its environment. Paint is cheap and can be applied easily. Paint is used to protect many everyday steel objects such as cars, trucks, trains, bikes, bridges etc. Paint also provides visual appeal.

2. Alloying:

The tendency of iron to oxidize can be greatly reduced by alloying it with other metals. For example, adding chromium to iron forms stainless steel, which is highly resistant to corrosion.

3. Coating with a thin layer of another metal:

Metals that readily corrode can be protected by coating with a thin layer of another metal that resists corrosion. This can be done by:

(a) Tinning

(b) Galvanizing

(c) Electroplating

(a) Tinning:

In the process of tin plating, clean iron sheet is dipped in a bath of molten tin. It is then passed through hot pair of rollers. Tin protects iron effectively, since, it is very stable.

(b) Galvanizing (Coating with Zinc):

The process of galvanizing consists of dipping a clean iron sheet in a hot zinc chloride bath and heating. After this sheet is rolled into zinc bath and cooled.

(c) Electroplating:

In electroplating an electrolytic process is used to deposit one metal on another metal.

4. Cathodic Protection:

Cathodic protection also called sacrificial protection is the process in which the metal to be protected from corrosion is made cathode and connected to an active metal such as magnesium or aluminum. These metals are more active than iron, they act as anode. The more active metal oxidizes itself and saves iron from corrosion. Cathodic protection is employed to prevent iron and steel structures such as pipes, tanks, oil rigs etc in the moist underground and marine environment. The large bars of magnesium are used to help protect the ship against rusting.

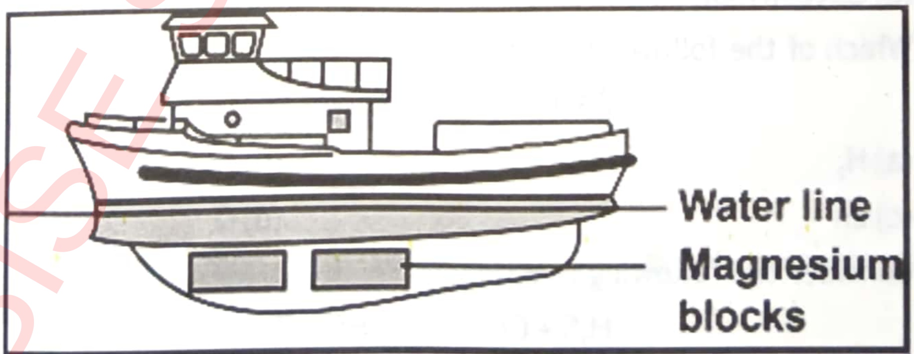


Figure 7.3: Cathodic Protection

KEY POINTS

- Oxidation is the gain of oxygen atom or loss of hydrogen atom or loss of electrons by a substance.
- Reduction is the loss of oxygen atom or gain of hydrogen atom or gain of electrons by a substance.
- Oxidation state or oxidation number is defined as the number of apparent charges that an atom will have in a molecule.
- The sum of oxidation state of all the atoms in a molecule of compound is zero.
- An oxidizing agent is the reactant containing the element that is reduced in a reaction.
- A reducing agent is the reactant containing the element that is oxidized in a reaction.
- Corrosion is the process in which a metal reacts with oxygen and moisture in the atmosphere.
- Electrolytic process used to deposit one metal on another metal is called electroplating.
- Cathodic protection is the process in which metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminum.

References for additional information

- B.Earl and LDR Wilford, Further Advanced Chemistry.
- B.Earl and LDR Wilford, Introducion to Advanced Chemistry.
- David E. Goldberg, Fundamental of Chemistry.
- Addison Wesley, Chemistry

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) In which of the following changes, the nitrogen atom is reduced.

(a) N_2 to NO (b) N_2 to NO_2

(c) N_2 to NH_3 (d) N_2 to HNO_3

(ii) Which of the following changes reaction is an example of oxidation

(a) Chlorine molecule to chloride ion (b) Silver atoms to silver (I) ion

(c) Oxygen molecule to oxide ion (d) Iron (III) ion to iron(II) ion

(iii) Which of the following elements in the given reaction is reduced?

$$ZnO + H_2 \longrightarrow Zn + H_2O$$

(a) H_2 (b) ZnO

(c) Zn (d) O

(iv) Consider the following reaction:

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

In this reaction H_2S behaves as

(a) Reducing agent (b) Oxidizing agent
(c) Catalyst (d) Electrolyte

(v) The oxidation state of Cr in $K_2Cr_2O_7$ is
(a) +12 (b) +6
(c) +3 (d) -6

2. Give short answer.

(i) What is oxidation state?
(ii) What is the oxidation number of Cr in chromic acid (H_2CrO_4)?
(iii) Identify reducing agent in the following reaction
(iv) $CuO + H_2 \longrightarrow Cu + H_2O$
(v) Why tin plated steel is used to make food cans?
(vi) Explain one example from daily life which involves oxidation-reduction reaction?

3. Compare and contrast oxidation and reduction.
4. Define oxidation and reduction in terms of loss or gain of electrons.
5. Explain how food and beverage industries deal with corrosion.
6. State the substances which are oxidized or reduced. Give reason for your answer.
(a) $N_2 + 3H_2 \longrightarrow NH_3$
(b) $CO_2 + 2Mg \longrightarrow 2MgO + C$
(c) $Mg + H_2O \longrightarrow MgO + H_2$
(d) $H_2S + Cl_2 \longrightarrow 2HCl + S$
(e) $2NH_3 + 3CuO \longrightarrow 3Cu + N_2 + 3H_2O$
7. (a) Define oxidation number or oxidation state.
(b) Find the oxidation state of nitrogen in the following compounds.
(i) NO_2 (ii) N_2O (iii) N_2O_3 (iv) HNO_3
8. Find the oxidation state of S in the following compounds.
(a) H_2S (b) H_2SO_3 (c) $Na_2S_2O_3$
9. (a) Define oxidizing and reducing agents.
(b) Identify the oxidizing agents and reducing agents in the following reactions:
(i) $H_2S + Cl_2 \longrightarrow 2HCl + S$
(ii) $2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$
(iii) $2KI + Cl_2 \longrightarrow 2KCl + I_2$
(iv) $Mg + 2HCl \longrightarrow MgCl_2 + H_2$
10. Hydrogen peroxide reacts with silver oxide and lead(II) sulphide according to the following equations.

Is hydrogen peroxide an oxidizing or reducing agent in these reactions. Argue.

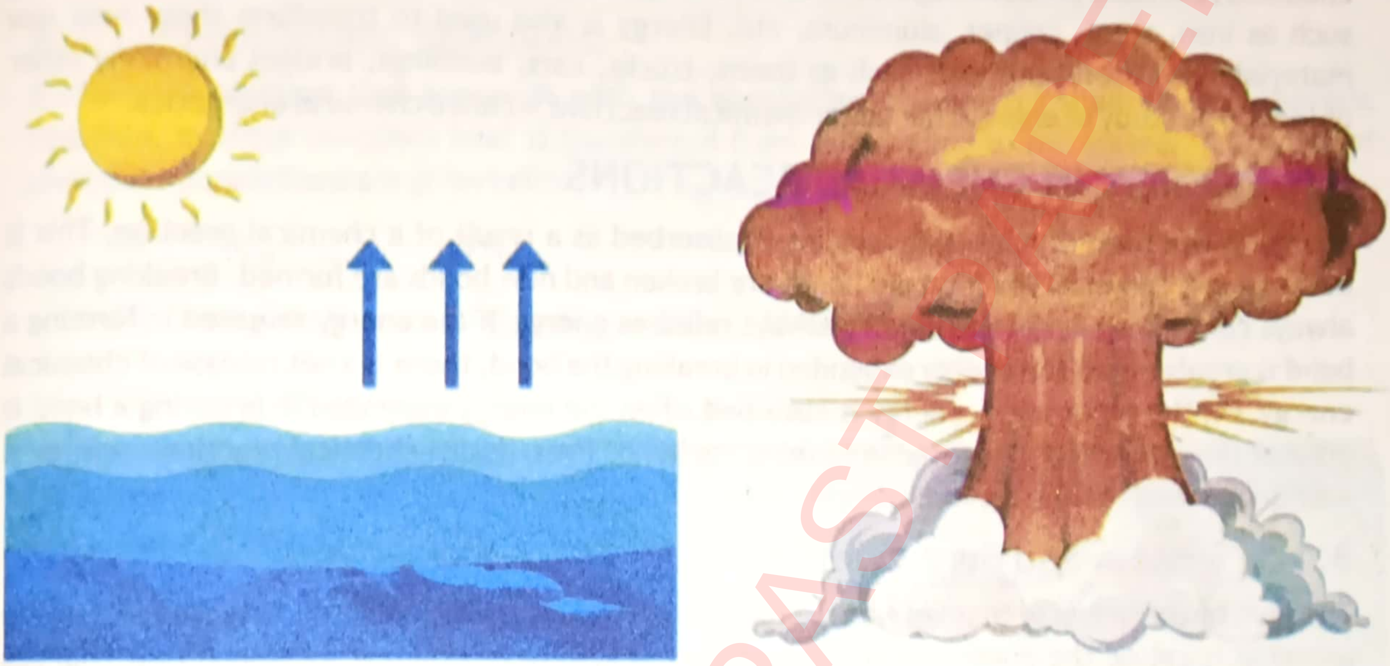
(i) $H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2$

(ii) $4H_2O_2 + PbS \longrightarrow PbSO_4 + 4H_2O$

PROJECT

Prepare a report about what type of chemical reaction is corrosion, giving suitable examples.

UNIT 08



ENERGETICS

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Explain the idea of a chemical system and its connections with its surroundings influences energy transfer during a chemical reaction.
- Differentiate between exothermic and endothermic reactions giving examples.
- State that thermal energy is called enthalpy change and recognize its sign as negative for exothermic and positive for endothermic reactions.
- Define activation energy as minimum energy that colliding molecules must have for a successful collision.
- Explain that activation energy depends on reaction pathway which can be changed using catalyst or enzyme (detailed pathways not required).
- Draw, label and interpret reaction pathway diagram for exothermic and endothermic reaction which includes enthalpy change, activation energy (uncatalyzed and catalyzed), reactants and products.
- Recognize that bond breaking is endothermic and bond making is exothermic processes.
- Explain that enthalpy change is sum of energies absorbed and released in bond breaking and bond forming.
- Calculate enthalpy change of a reaction from given bond energy values.
- Explain how respiration(aerobic and anaerobic), an exothermic process, provides energy for biological systems and lipids as reserve stores of energy.

Introduction

Every process in this universe, whether it is in living cells, test tubes, atmosphere or water, etc., involves a change in energy. Some processes release energy, others require energy. Many chemical reactions produce huge amounts of energy, which is used to produce new raw materials such as iron, steel, copper, aluminum, etc. Energy is also used to transform these new raw materials into useful products such as trains, trucks, cars, buildings, bridges and many other objects. The study of energy changes in chemical reactions is called chemical energetics.

8.1 ENERGY IN CHEMICAL REACTIONS

Energy in the form of heat is developed or absorbed as a result of a chemical reaction. This is because in a chemical reaction old bonds are broken and new bonds are formed. Breaking bonds always consumes energy and binding always releases energy. If the energy released in forming a bond is greater than the energy expended in breaking the bond, there is a net release of chemical energy. On the other hand, energy is absorbed when the energy expended in breaking a bond is greater than the energy released in forming the bond. Thus, during chemical reactions, energy is exchanged with the surroundings.

8.1.1 System and Surroundings

The part of the universe that we want to focus our attention on is called a system. The rest of the universe is called the environment. In chemistry, a system is usually a substance that changes physically or chemically. For example, when studying the reaction of limestone and hydrochloric acid solution in a test tube, limestone and hydrochloric acid solution form a system. The test tube and everything around the test tube is the environment. Similarly, when studying the thermal decomposition of a compound, the sample of the compound would be the system. While the beaker, heat source, and everything else would be the environment.

8.2 THERMOCHEMICAL REACTIONS

When a chemical change takes place, energy is exchanged between system and its surroundings. Energy has many forms such as heat, light, work etc. A chemical reaction which proceeds with the evolution or absorption of heat is called a thermochemical reaction. **A balanced chemical equation which also shows heat change of a chemical reaction is called thermochemical equation. The branch of chemistry which deals with the heat or thermal energy changes in chemical reactions is called thermochemistry.**

For example

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H^{o} = -393.5\,kJ$$

There are two types of thermochemical reactions.

8.2.1 Exothermic Reactions

A chemical reaction that proceeds with the evolution of heat is called an exothermic reaction. In an exothermic reaction the chemical system transfers energy to the surroundings as the reactants are converted to products e.g. burning of fuels is a highly exothermic reaction. The energy released can be used to heat a room, or to drive an engine or to cook food. Examples of exothermic reactions are given below:

(i) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H^\circ = -393.5kJ$

(ii) $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$ $\Delta H^\circ = -571.6kJ$

(iii) $C_{(s)} + \frac{1}{2} O_{2(g)} \longrightarrow CO_{(g)}$ $\Delta H^\circ = -110.5kJ$

8.2.2 Endothermic Reactions:

A chemical reaction that proceeds with the absorption of heat is called an endothermic reaction. In these reactions heat is transferred from surrounding to the system. Examples of endothermic reactions are given below:

(i) $H_{2(g)} + I_{2(g)} \longrightarrow 2HI_{(g)}$ $\Delta H^\circ = +53.8kJ$

(ii) $C_{(s)} + H_2O_{(g)} \longrightarrow CO_{(g)} + H_{2(g)}$ $\Delta H^\circ = +131.4kJ$

(iii) $N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)}$ $\Delta H^\circ = +180.5kJ$

CONCEPT ASSESSMENT EXERCISE 8.1

Classify the following processes as exothermic or endothermic.

(a) Freezing of water
(b) Combustion of methane
(c) Sublimation of dry ice
(d) $H_2O_{(g)} \longrightarrow H_2O_{(l)}$
(e) decomposition of limestone.

8.3 ENTHALPY OF REACTION

The amount of heat or thermal energy evolved or absorbed in a chemical reaction is called enthalpy of reaction. Its sign is negative for exothermic and positive for endothermic reactions.

Enthalpy of reaction measured at 25°C (or 298K) and one atmospheric pressure is known as standard enthalpy change. It is denoted by ΔH°.

(i) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H^\circ = -393.5k\ J$

(ii) $H_{2(g)} + I_{2(g)} \longrightarrow 2HI_{(g)}$ $\Delta H^\circ = +53.8kJ$

Which of the above reaction is endothermic?

8.4 BOND ENERGY AND BOND DISSOCIATION ENERGY

When a chemical reaction occurs, old bonds are broken and new bonds are formed. Breaking bonds always requires energy, and forming a bond always releases energy. The amount of energy required to break one mole of a particular bond to form neutral atoms is called the bond dissociation energy. In contrast, the amount of energy released when neutral atoms form one mole of a bond is called bond energy. The difference between the bond dissociation energy and the bond energy determines whether the reaction absorbs or releases energy.

The enthalpy change in a chemical reaction is the sum of energies absorbed and released in bond breaking and bond forming.

ΔH° = Sum of bond dissociation energies of reactants − Sum of bond energies of products

Example 8.1: Calculate the enthalpy of the reaction between hydrogen and iodine to form hydrogen iodide from the given bond energy data. Bond energy of H–H, I–I, H–I bonds are 436kJ/mol, 151kJ/mol and 299kJ/mol respectively

Problem solving strategy:

1. Write the balanced chemical equation.
2. Show all the reactant and the products in the gaseous state.
3. Substitute the relevant bond energy values in the formula and solve.

ΔH° = Sum of bond dissociation energies of reactants − Sum of bond energies of products

Solution:

$$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)} \quad \Delta H^\circ = ?$$

$$\Delta H^\circ = [\text{B.E of H–H} + \text{B.E. of I–I}] - [2 \times \text{B.E of H–I}]$$

$$= [436 + 151] - [2 \times 299]$$

$$= 587 - 598$$

$$= -11 \text{ kJ/mol}$$

Note that the enthalpy of reaction calculated using bond energy data are often different from values determined experimentally.

CONCEPT ASSESSMENT EXERCISE 8.2

Example: Calculate the enthalpy of the following reaction from the given bond energy data. Bond energy of H–H, F–F, H–F bonds are 436kJ/mol, 155kJ/mol and 567kJ/mol respectively.

8.5 ACTIVATION ENERGY

Chemical reactions involve the breaking and forming of chemical bonds. These changes are accompanied by changes in energies. Collision theory was proposed to explain the observed reaction kinetics. For a chemical reaction to occur, the bonding atoms or molecules must collide with each other. These collisions can be effective or ineffective depending on the energy and direction of the colliding particles. An effective collision can only occur if the energy of the colliding particles is high enough to overcome the repulsion between the electrons around the reacting particles. The correct orientation means that at the moment of collision, the atoms needed to form new bonds must collide with each other. The minimum amount of energy that, in addition to the average kinetic energy, particles must have an effective collisions is called the activation energy. No reaction occurs if the energy of the reacting particles is lower than the activation energy. Thus, the speed of a reaction depends on

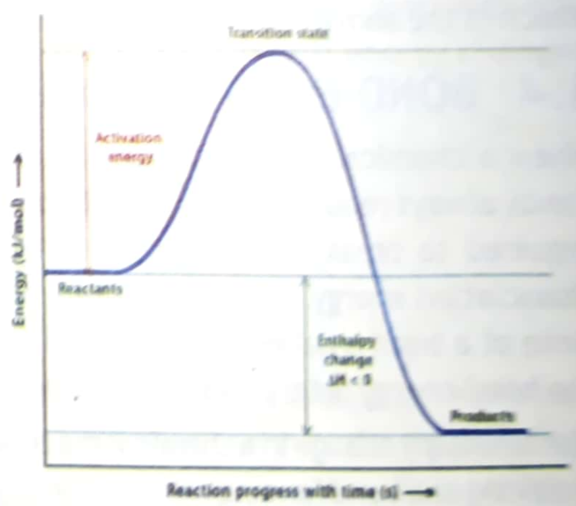
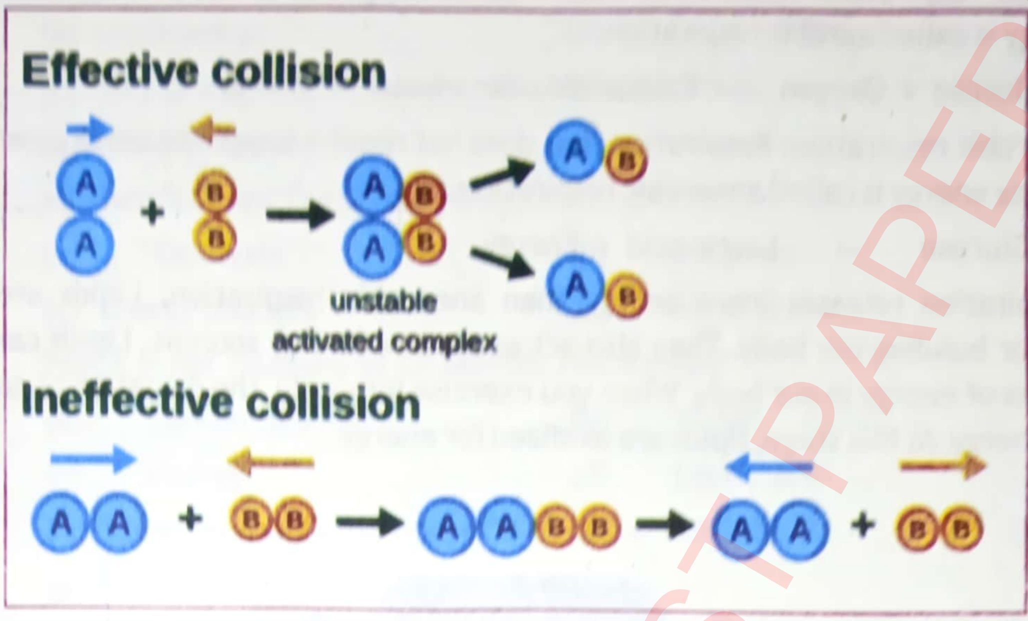


Figure 8.1: Activation energy

its activation energy. The higher the activation energy, the lower the reaction rate. For example, the reaction between A_2 and B_2 molecules.



8.6 CATALYST

Many industrial reactions are carried out at high temperatures over a period of time to maximize the amount of product that can be synthesized. High temperature reactions cause safety problems and many chemicals are not stable at high temperatures. So it would be useful to use another method to increase the speed of chemical reactions. Another way to increase the reaction rate is to change this mechanism in a way that lowers the activation energy. This can be done by adding a catalyst. **A substance that accelerates a chemical reaction, but remains chemically unchanged at the end of the reaction, is called a catalyst**, and the phenomenon is called catalysis. The catalyst provides a new mechanism for the reaction with low activation energy (Figure 8.3). Thus, a catalyst increases the rate of a reaction by lowering its activation energy. The catalyst does not affect the overall thermodynamics or enthalpy of the reaction.

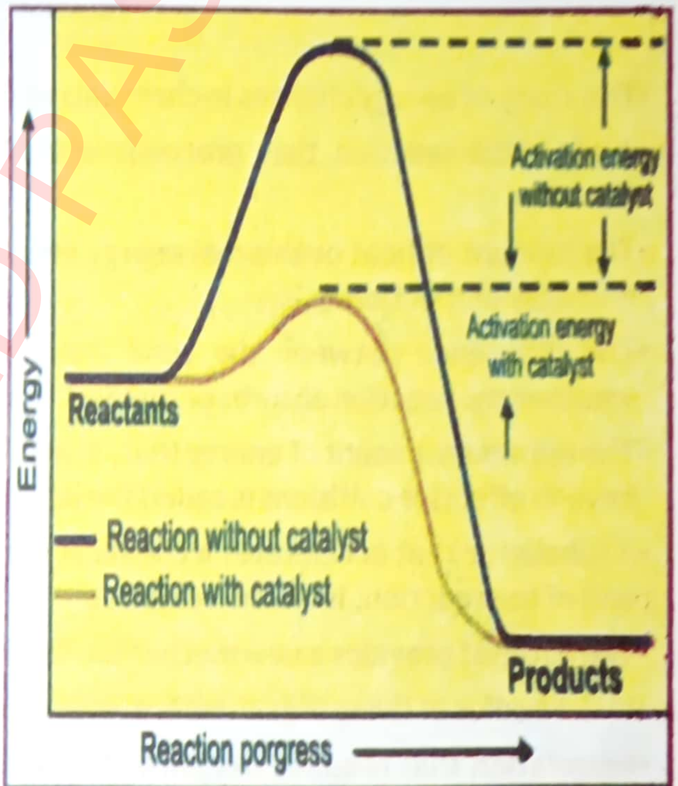


Figure 8.3: Effect of Catalyst on energy of activation

In the bodies of living organisms enzymes catalyze chemical reactions.

8.7 RESPIRATION

Where do we get energy for our body? Respiration is a biochemical process in which energy is released from food in a biological system. During this process, glucose is oxidized in the body of living organisms and energy is released. Therefore, respiration is an exothermic reaction. There are two types of respiration processes.

1. Aerobic respiration: Respiration that requires oxygen to break down glucose to release energy is called aerobic respiration.

   Glucose + Oxygen → Carbon dioxide + water + Energy

2. Anaerobic respiration: Respiration that does not require oxygen to break down glucose to release energy is called anaerobic respiration.

   Glucose → Lactic acid + Energy

Aerobic respiration releases more energy than anaerobic respiration. Lipids are important substances for building our body. They also act as reserve energy sources. Lipids can store very large amounts of energy in our body. When you exercise intensely, the oxidation of glucose is not enough for energy. At this stage, lipids are oxidized for energy.

KEY POINTS

- The study of energy changes in chemical reactions is called chemical energetics.
- A chemical reaction that proceeds with the evolution of heat is called an exothermic reaction.
- The amount of heat or thermal energy evolved or absorbed in a chemical reaction is called enthalpy of reaction.
- The difference between the bond dissociation energy and the bond energy determines whether the reaction absorbs or releases energy.
- The minimum amount of energy that, in addition to the average kinetic energy, particles must have in effective collisions is called the activation energy.
- A substance that accelerates a chemical reaction, but remains chemically unchanged at the end of the reaction, is called a catalyst.
- The catalyst provides a new mechanism for the reaction with low activation energy.
- In the bodies of living organisms enzymes catalyze chemical reactions.
- Respiration that requires oxygen to break down glucose to release energy is called aerobic respiration.
- Respiration that does not require oxygen to break down glucose to release energy is called anaerobic respiration.
- Lipids acts as reserve energy sources.

References for additional information

- Zumdahl, Introductory Chemistry.
- Raymond Chang, Essential Chemistry.

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) If the ΔH value is negative than reaction will be
(a) Exothermic (b) Endothermic
(c) May or may not be Exothermic or Endothermic
(d) None of these

(ii) All chemical reactions involve
(a) Catalysts (b) Enzymes
(c) Energy changes (d) All of these

(iii) Which is not released in an aerobic respiration?
(a) Carbon dioxide (b) Water
(c) Energy (d) Lactic acid

(iv) A catalyst increases the rate of a chemical reaction by
(a) increasing activation energy
(b) increasing the enthalpy of reaction
(c) Decreasing the enthalpy of reaction
(d) None of these

(v) Activation energy of a chemical reaction must be __________ the average kinetic energy of reacting molecules
(a) Lower than (b) greater than
(c) equal to (d) None of these

2. **Give short answer.**

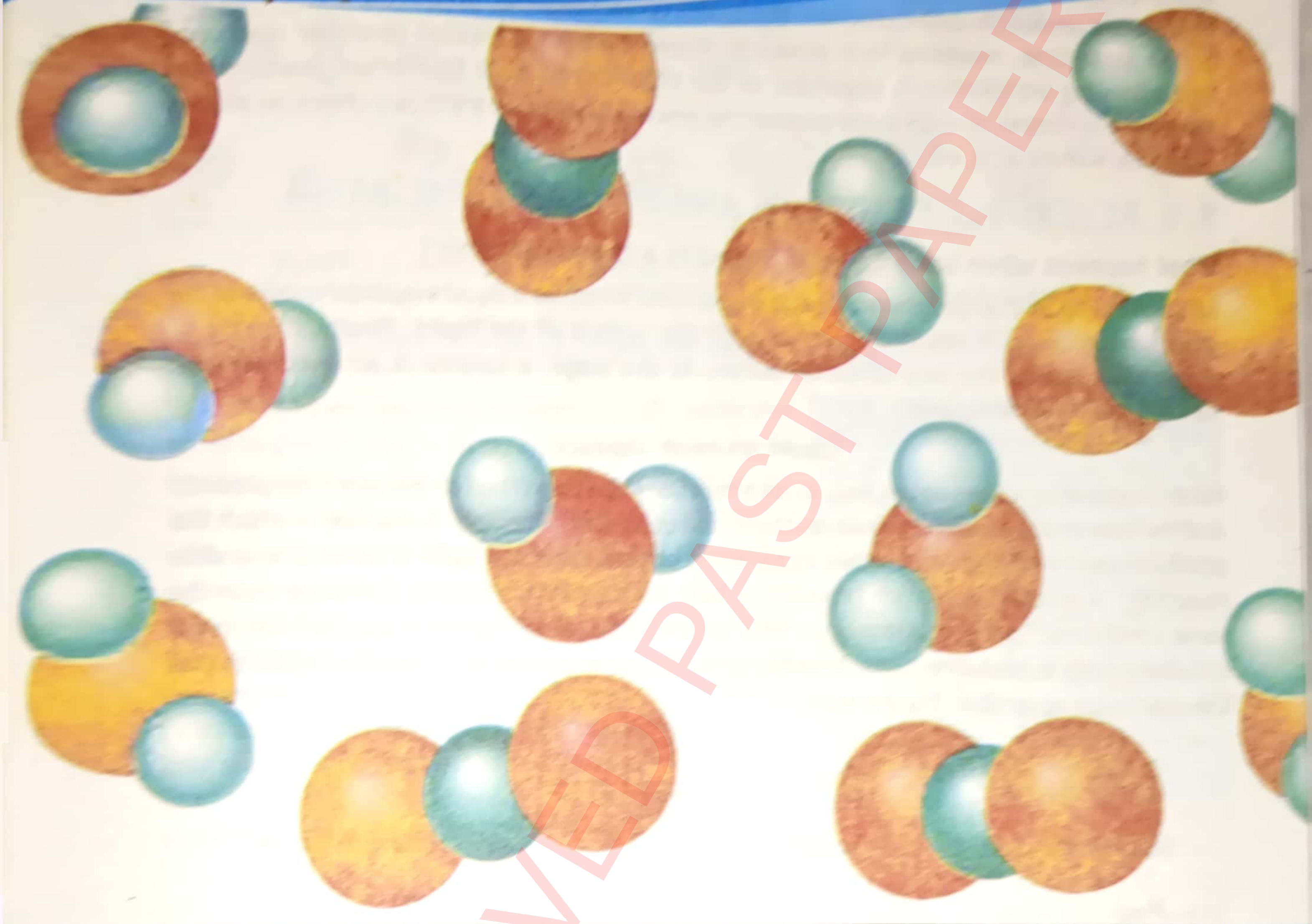
(i) Define exothermic and endothermic reactions.
(ii) Define enthalpy of a chemical reaction.
(iii) What is anaerobic respiration?
(iv) Define activation energy.
(v) What is the role of a catalyst in a chemical reaction.
(vi) Differentiate between aerobic and anaerobic respiration.

3. How can you determine the enthalpy of a chemical reaction?
4. Explain, how does the process of respiration provides us energy?
5. Draw labeled reaction pathway diagram for an exothermic and an endothermic reaction.
6. Calculate the enthalpy of reaction between hydrogen and chlorine to form hydrogen chloride from the given bond energy data. Bond energy of H-H, Cl-Cl, H-Cl are 436kJ/mol, 243kJ/mol and 432kJ/mol respectively.
7. Justify the statement that the process of respiration is crucial for us.

PROJECT

Create a chart showing pathway diagram for exothermic and endothermic reactions, which includes enthalpy change, activation energy (catalyzed and uncatalyzed), reactants and products.

UNIT 09



CHEMICAL EQUILIBRIUM

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Recognize that reversible reaction are shown by $\rightleftharpoons$ and may not go to completion.
- Describe how changing the physical conditions of a chemical equilibrium system can redirect reversible reactions (a) effect of heat on hydrated compounds (b) addition of water to anhydrous substances in particular copper (II) sulphate and cobalt (II) chloride.
- State that reversible reactions can achieve equilibrium in a closed system when rate of forward and backward reactions are equal

INTRODUCTION:

A complete reaction is a reaction in which all reactants have been converted to products. However, many important chemical reactions are not completed and a mixture of products and reactants is formed. In such a reaction, the products react together to form the reactants again. At the same time, reactants form products. These reactions are called reversible reactions. Understanding equilibrium is important in the chemical industry. Equilibrium reactions are involved in a number of steps in the commercial production of many important chemicals such as ammonia, sulfuric acid, etc.

9.1 REVERSIBLE REACTIONS AND DYNAMIC EQUILIBRIUM.

What happens when some liquid is placed in a closed container?

Some of the liquid changes physically as it evaporates. When more liquid evaporates, some of the vapor condenses as a result of collision with the surface of the liquid. Finally, the rate of evaporation equals the rate of condensation. At this stage, a balance is achieved between forward and backward changes.

$$\text{Liquid} \rightleftharpoons \text{Vapours}$$

Many chemical reactions do not end. In such reactions, the conversion of reactants into products and the conversion of products into reactants can occur simultaneously. **A reaction in which the products can react with each other to form the original reactants again is called a reversible reaction.** A reversible reaction proceeds in both the forward and reverse directions under the same conditions. These reactions never end. All reversible changes (physical and chemical) occur simultaneously in both directions. A double arrow $\rightleftharpoons$ in a chemical equation indicates that the reaction is reversible. For example:

$$\underset{\text{Sulphur dioxide}}{2SO_{2(g)}} + \underset{\text{Oxygen}}{O_{2(g)}} \underset{450^\circ C,\ 200\ atm}{\overset{V_2O_5}{\rightleftharpoons}} \underset{\text{Sulphur trioxide}}{2SO_{3(g)}}$$

$$\underset{\text{Nitrogen}}{N_{2(g)}} + \underset{\text{Hydrogen}}{3H_{2(g)}} \underset{400^\circ C,\ 200\ atm}{\overset{Fe}{\rightleftharpoons}} \underset{\text{Ammonia}}{2NH_{3(g)}}$$

$$\underset{\text{Nitrogen}}{N_{2(g)}} + \underset{\text{Oxygen}}{2O_{2(g)}} \underset{\text{spark}}{\overset{\text{Electric}}{\rightleftharpoons}} \underset{\text{Nitrogen dioxide}}{2NO_{2(g)}}$$

Consider what happens when SO_2 and O_2 gases are mixed in a sealed container (Figure 9.1)

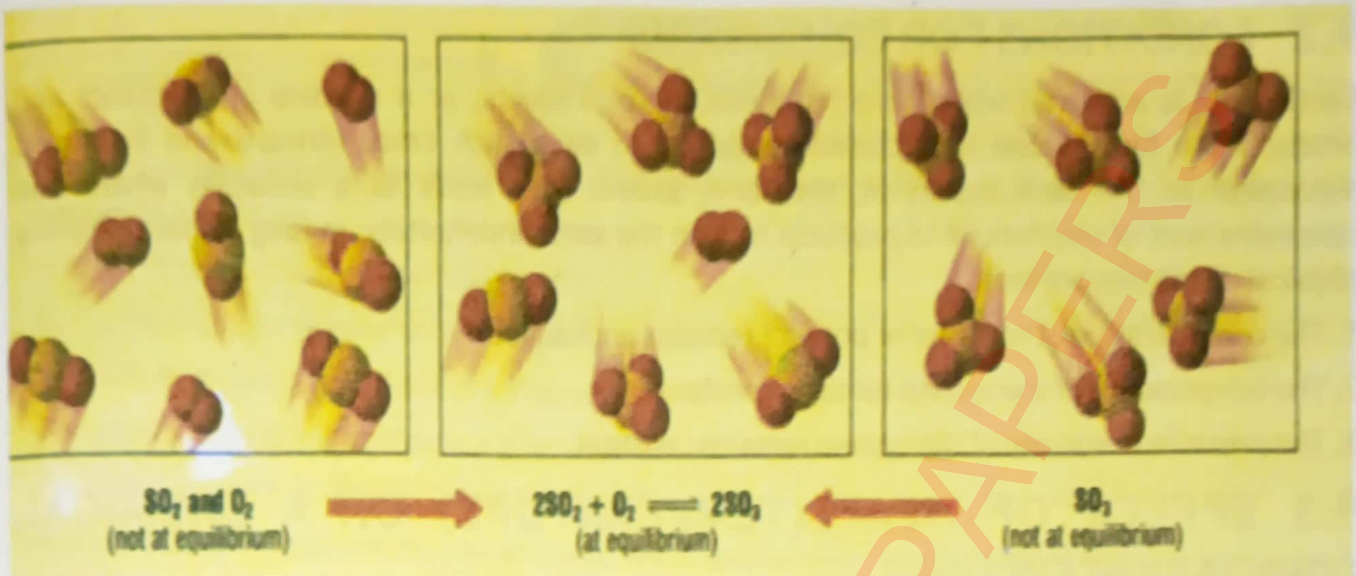


Figure 9.1: Reaction between SO_2 and O_2

So$_2$ and O_2 molecules react to form SO_3. SO_3 molecules break to give SO_2 and O_2. What types of molecules are in equilibrium? In the first reaction (from left to right), SO_2 and O_2 produce SO_3. In the second reaction (from right to left), SO_3 decomposes into SO_2 and O_2. Which reaction is called a forward reaction? Which reaction is called the reverse reaction? In the beginning, there is no SO_3. So the rate of the reverse reaction is zero. Due to the high concentration of the reactants, the speed of the forward reaction is the highest. As the reaction progresses, the concentration of the reactants gradually decreases and the speed of the forward reaction decreases accordingly. (figure)

Science Titbits

When fizzy drinks are made, CO_2 is dissolved in the liquid drink under pressure and sealed. When you remove lid of the bottle, bubbles CO_2 of suddenly appear. When you put the lid back on the bottle, the bubbles stop. This is due to the following equilibrium.

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)}$$

The forward reaction happens during manufacturing and the reverse reaction happens on opening

$$2SO_{2(g)} + O_{2(g)} \longrightarrow SO_{3(g)}$$

As the concentration of SO_3 increases, a small amount of SO_3 slowly decomposes into SO_2 and O_2. This means that the reverse reaction has started. In this reaction, SO_3 acts as a reactant and produces SO_2 and O_2. That is, the opposite reaction

$$2SO_{3(g)} \longrightarrow 2SO_{2(g)} + O_{2(g)}$$

As the concentration of SO_3 increases, the reverse reaction accelerates. In the end, the two rates are equal. At this point, SO_3 decomposes into SO_2 and O_2 as fast as SO_2 and O_2 produce SO_3. At this point, the reaction has reached equilibrium. (Figure 9.1) **The state of a chemical reaction where the forward and reverse reactions occur at the same rate is called chemical equilibrium.** Chemical equilibrium is dynamic equilibrium. This is because reactions do not stop when they reach equilibrium. Individual molecules are constantly reacting. However, the actual quantities of reactants and products do not change. This means that the concentration of reactants and products becomes constant in the equilibrium state.

9.2 CONDITIONS FOR EQUILIBRIUM

Equilibrium is achieved when pure reactants, pure products, or a mixture of reactants and products are first placed in a closed container. In each such case, forward and backward movement in the tank occurs at the same speed. This leads to a situation where the concentrations of reactants and products remain the same indefinitely, as long as the following physical conditions are met:

1. The concentration of reactant or product remains unchanged.
2. The temperature of the system remains constant.
3. The pressure or volume of the system remains constant.

9.3 EFFECT OF HEAT (TEMPERATURE) ON A CHEMICAL EQUILIBRIUM SYSTEM

Activity:

Materials required Test tube, dropper, heating system, hydrated copper(II)sulphate.

Procedure:

- Place 5 g of hydrated copper(II) sulphate in a test tube. and heat slowly.
- Observe the colour change from blue to white.
- Allow the test tube and its contents to cool to room temperature.
- Add a few drops of water to the test tube using a dropper.
- Observe the colour change from white to blue again. When copper (II) sulphate is heated, the water in it is removed, forming anhydrous copper (II) sulphate, which is a white solid. This copper (II) sulphate changes back to the hydrated form on adding water.

NOTE:

Note: Copper sulphate is a harmful and toxic compound, so handle it with care. Wear safety glasses and gloves. Do this task in the presence of your teacher.

$$CuSO_4.5H_2O \text{ (blue solid)} \rightleftharpoons CuSO_4\text{(white solid)} + 5H_2O$$

Likewise

Hydrated cobalt(II) chloride is a pink solid. When heated, it loses water and becomes anhydrous cobalt(II) chloride, a blue solid. So the equilibrium shifts towards right. But when water is added to it, it absorbs water and the equilibrium shifts to the left to form hydrated cobalt(II) chloride again

$$CoCl_2.6H_2O\text{(pink solid)} \rightleftharpoons CoCl_2\text{(blue solid)} + 6H_2O$$

KEY POINTS

- A reaction in which the products can react together to re-form the original reactants is called reversible reaction.
- A reversible reaction is shown by symbol .
- Anhydrous copper(II)sulphate is a white solid.
- Hydrated copper(II)sulphate is a blue solid.
- Anhydrous cobalt(II)chloride is a blue solid.
- Hydrated cobalt(II)chloride is a pink solid.
- A state of a chemical reaction in which forward and reverse reactions take place at the same rate is called chemical equilibrium.

References for additional information

- Chemistry, Roger Norris, Lawrie Ryen and David Acaster.
- Principals of chemical equilibrium, Kenneth Denbigh.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which is true about the equilibrium state?

(a) The forward reaction stops.

(b) The reverse reaction stops.

(c) Both forward and reverse reactions stop.

(d) Both forward and reverse reactions continue at the same rate.

(ii) When a mixture of and is sealed in a flask and temperature is kept at 25°C, following equilibrium is established.

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

Which substance or substances will be present in the equilibrium mixture?

(a) H_2 and I_2

(b) HI only

(c) H_2 only

(d) H_2 , I_2 and HI

(iii) Concentration of reactants and products at equilibrium remains unchanged if

(a) concentration of any reactant or product is not changed.

(b) temperature of the reaction is not changed.

(c) pressure or volume of the system is not changed.

(d) all of the above are observed

(iv) Which of the following does not happen, when a system is at equilibrium state?

(a) forward and reverse reactions stop.

(b) forward and reverse rates become equal.

(c) concentration of reactants and products stop changing.

(d) reaction continues to occur in both the directions.

(v) In an irreversible reaction equilibrium is

(a) established quickly

(b) established slowly

(c) never established

(d) established when reaction stops.

2. Give short answer.

(i) Differentiate between forward and reverse reactions.

(ii) What is chemical equilibrium?

(iii) Write two chemical equations of reversible reactions.

(iv) Write down the conditions for equilibrium.

3. Coal reacts with hot steam to form CO and H_2. These substances react further in the presence of a catalyst to give methane and water vapour.

$$CO_{(g)} + 3H_{2(g)} \rightleftharpoons CH_{4(g)} + H_2O_{(g)}$$

Write forward and reverse reactions for it.

4. How does temperature affect cobalt chloride equilibrium?

THINK TANK

1. **Bromine chloride (BrCl) decomposes to form chlorine and bromine. Write reversible chemical reaction for this reaction.**

UNIT 10

ACIDS, BASES AND SALTS

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define Bronsted-Lowry acids as proton donors and Bronsted-Lowery bases as proton acceptors.
- Recognize that aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions.
- Define a strong acid and a base as an acid or base that completely dissociates in aqueous solution and weak acids and bases that partially dissociates in aqueous solution.(some example include: Students writing symbol equations to show these for hydrochloric acid, sulphuric acid, nitric acid, and ethanoic acid).
- Formulate dissociation equations for an acid or base in aqueous solution.
- Recognize that bases are oxides or hydroxides of metals and that alkalis are water-soluble bases.
- Describe the characteristic properties of acids in terms of their reactions with metals, bases, and carbonates.
- Identify the characteristic properties of bases in terms of their reactions with acids and ammonium salts.
- Define acid rain.
- Discuss effects of acid rain and relate them with the properties of acids.

INTRODUCTION

You often use acids and bases in all areas of life. For example, vinegar, aspirin, lemon juice, cola drinks, apple, tomato and toilet cleaners contain acids. Substances containing bases such as drain cleaner, antacid tablets, baking powder, soda, etc. You eat and drink certain acids and bases and your body produces them. From "acid indigestion" to "acid rain," the word acid appears frequently in news and advertisements. What is acid rain? This chapter will help you understand which substances are called acids and which are bases. How are they classified? What happens when an acid reacts with a base? Why do we use lemon juice on fish? In this chapter, you will learn about the chemistry of acids and bases. This will help you better understand these important classes of compounds. What do we mean by the pH of a solution like acid rain? Acids are widely used in the manufacture of fertilizers and in the food industry.

10.1 CONCEPTS OF ACIDS AND BASES

Acids and bases are generally recognized by their characteristic properties. Table 10.1 shows such properties.

Table 10.1: Some characteristic properties of acids and bases

Sr. No.	Property	Acid	Base
1	Taste	Sour	Bitter
2	Effect on blue litmus	Turns red	No effect
3	Effect on red litmus	No effect	Turns blue
4	Effect on skin	Corrosive	Corrosive
5	Electrical conductivity	Aqueous solutions conduct electricity	Aqueous solutions conduct electricity

10.1.1 Arrhenius Concept of Acids and Bases

In 1887, a Swedish chemist Svante Arrhenius proposed the first successful theory of acids and bases. According to this theory

An acid is a substance that ionizes in water to produce H^+ ion sand a base is a substance that ionizes in water to produces OH^- ions.

For example,

$$HCl_{(g)} \xrightleftharpoons{H_2O} H^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$NaOH_{(s)} \xrightleftharpoons{H_2O} Na^+{}_{(aq)} + OH^-{}_{(aq)}$$

Which substances in the following reactions are acids or bases?

$$HNO_{3\,(l)} \xrightleftharpoons{H_2O} H^+{}_{(aq)} + NO_3^{-1}{}_{(aq)}$$

$$H_2SO_{4\,(l)} \xrightleftharpoons{H_2O} 2H^+{}_{(aq)} + SO_4^{-2}{}_{(aq)}$$

$$KOH_{(s)} \xrightleftharpoons{H_2O} K^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$NH_4OH_{(aq)} \xrightleftharpoons{H_2O} NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Table 10.2 shows some common acids and table 10.3 shows some common bases.

Table 10.2 Some Common Acids

Name	Formula	Common use
Hydrochloric acid	HCl	Cleaning of metals, bricks and removing scale from boilers
Nitric acid	HNO_3	Manufacture of fertilizers, explosives
Sulphuric acid	H_2SO_4	Manufacture of many chemicals, drugs, dyes, paints and explosives.
Phosphoric acid	H_3PO_4	Manufacture of fertilizers, acidulant for food

Table 10.3 Some Common Bases

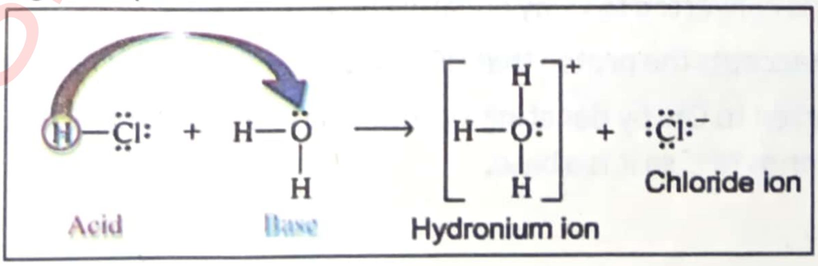
Name	Formula	Common use
Sodium hydroxide	$NaOH$	Soap making, drain cleaners
Potassium hydroxide	KOH	Making liquid soap, shaving cream
Calcium hydroxide	$Ca\,(OH)_2$	Making mortar, plasters, cement
Magnesium hydroxide	$Mg(OH)_2$	Antacid, laxative

10.1.2 The Bronsted-Lowery Concept of Acids and Bases

The Arrhenius theory has its limitations. This applies to aqueous solutions. This does not explain why compounds like CO_2, SO_2, etc. are acidic. Why are substances like NH_3 bases? There is no H in CO_2 and no O in NH_3.

In 1923, J. N. Bronsted and T. M. Lowery independently proposed another theory to overcome the shortcomings of the Arrhenius theory. This theory is known as the Bronsted-Lowery theory. According to this theory, an acid is a proton donor and a base is a proton acceptor.

Consider the following example



Q. 1. Whichsubstance is donating proton?

____

Q. 2. Which substance is accepting proton?

____

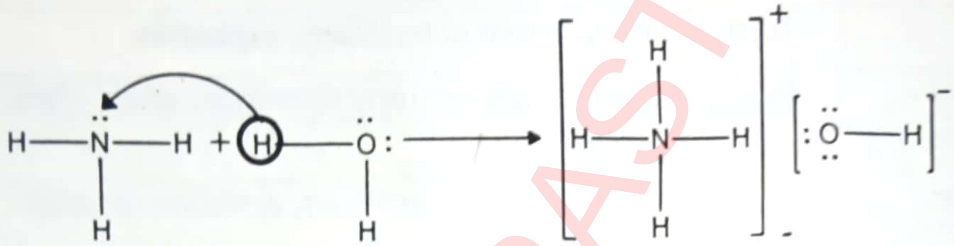
Q. 3. Which substance is acid?

____

Q. 4. Which substance is base?

____

Where does OH^- come from when ionizing bases such as ammonia? The Arrhenius theory is not sufficient to answer this question, but the Bronsted-Lowry theory explains how ammonia acts as a base in water. Ammonia is a gas at room temperature. When it is dissolved in water, the following reaction occurs.



Which substance is donating proton, NH_3 or H_2O? Which substance is proton acceptor? All the acids included in the Arrhenius Theory are also acids in the Bronsted-Lowery Theory. However, all the bases included in Bronsted-Lowery theory except OH^- are not Arrhenius bases. Consider above two examples. In one example, water molecule accepts a proton and in the other water donates a proton. This means water behaves like an acid as well as a base. It is amphoteric in nature. Substances that react with both acids and bases are called amphoteric substances.

Example 10.1: Classify substances as acids or bases or as proton donor or proton acceptor

Identify Bronsted-Lowery acids or bases in the following reactions.

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
$$H_2O + NH_3 \longrightarrow NH_4^+ + OH^-$$

Problem solving strategy:

1. An acid is a proton donor. After donating proton, an acid forms a negative ion.
2. A base is a proton acceptor. After accepting proton from an acid it forms a positive ion.

Solution:

1. Because HCl is converted to Cl^- by donating proton, HCl is an acid.
2. Because H_2O accepts the proton that HCl donates and forms H_3O^+, wateris a base.
3. H_2O is converted to OH^- by donating a proton, so H_2O is an acid. Because NH_3 accepts the proton and forms NH_4^+ so it is a base.

CONCEPT ASSESSMENT EXERCISE 10.1

Identify Bronsted acids and Bronsted bases in the following reactions.

1. $H_2SO_4 + H_2O \rightleftharpoons HSO_4^- + H_3O^+$
2. $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
3. $H_2S + NH_3 \rightleftharpoons NH_4^+ + HS^-$

10.2 STRENGETH OF ACIDS AND BASES

All the acids and bases do not ionize in aqueous solutions to the same degree. Therefore, they have different strengths.

10.2.1 Strong and weak acids

An acid that ionizes completely in aqueous solution is called a strong acid. For example, HCl, HNO_3, H_2SO_4 etc are strong acids. They ionize 100% in aqueous solution. All the molecules of strong acids ionizze in water.

$$HCl_{(g)} \xrightarrow{H_2O} H^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$HNO_{3\,(l)} \xrightarrow{H_2O} H^+{}_{(aq)} + NO_3^{1-}{}_{(aq)}$$

$$H_2SO_{4\,(l)} \xrightarrow{H_2O} 2H^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Acids that do not ionize completely in aqueous solutions are called weak acids. Fewer molecules of weak acids ionize in water. For example, ethanioc acid(acetic acid) which is found in vinegar ionizes only up to 5% in water. So, etanoic acid is a weak acid.

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

10.2.2 Strong and weak bases

Like acids, bases can also ionize in water to different degree. A base that ionizes completely or 100% in aqueous solution is termed as strong base. For example, NaOH, KOH, $Ca(OH)_2$ etc are strong bases.

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$KOH_{(s)} \xrightarrow{H_2O} K^+{}_{(aq)} + OH^-{}_{(aq)}$$

A base that ionizes to a little extent is called a weak base.Such bases produce fewer OH^- ions in aqueous solution. $Al(OH)_3$, NH_3 etc are weak bases.

$$H_2O + NH_3 \longrightarrow NH_4{}^+ + OH^-$$

Alkalis

A base which is soluble in water is called an alkali. This means that all the bases are not alkalis. On the other hand all the alkalis are bases. Which is an alkali KOH or NaOH?

Many bases do not dissolve in water. For example, copper hydroxide($Cu(OH)_2$), aluminium hydroxide$Al(OH)_3$ and ferric hydroxide$Fe(OH)_3$. All hydroxides are bases and only water soluble bases are alkalis. Many household items such as soaps, detergents, shampoos, toothpaste contain alkali.

10.3 CHARACTERISTIC PROPERTIES OF ACIDS

General properties of acids are as follows:

1. Acids have sour raste.
2. Acids change the colour of blue litmus paper to red.
3. Acids react with most metals and corrode them. Acids combine with metals like zinc, magnesium, aluminium and calcium to form their salts and hydrogen. The hydrogen gas is liberated in the form of bubbles. Zinc reacts with hydrochloric acid to produce zinc chloride with the leberation of hydrogen gas. Similarly, magnesium reacts with sulphuric acid to produce magnesium sulphate and hydrogen. The reaction of metals with acids can be described by the following equation.

   Metal + Acid → Salt + Hydrogen

   $Zn + HCl \rightarrow ZnCl_2 + H_2$

   $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

4. Acids react with metal carbonates to form salts, water with the liberation of carbon dioxide. The liberated carbon dioxide forms bubbles in water. For example hydrochloric acid reacts with sodium carbonate to forn sodiun chloride, carbon dioxide and water. Similarly, sulphuric acid reacrs with calcium carbonate(lime stone or marble) to produce calcium sulphate, carbon dioxide and water. The reaction of metal carbonate with acids can be represented by the following general reaction.

   Metal carbonate + Acid → Salt + Water + Carbon dioxide

   $Na_2CO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

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

   This reaction is used in the industrial preparation of glass, paper and soap.

5. Acids combine with bases to produce salt and water. This reaction is called as neutralization reaction. For example hydrochloric acid neuutralizes sodium hydroxide to form sodium chloride and water. Similarly sulphuric acid combines with potassium hydroxide to produce potassium sulphate and water. Neutralization reaction can be represented by the following general reaction.

   Acid + Base → Salt + Water

   $HCl + NaOH \rightarrow NaCl + H_2O$

   $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$

Normal rain is slightly acidic due to dissolved carbon dioxide. As a result of human activity, many oxides enter the atmosphere, which makes rainwater more acidic, which falls as acid rain .Acid rain contains dissolved nitric and sulphuric acid. Due to the corrosive nature of acids, acid rain can damage structures, buildings, and statues containing metals and metal carbonates. You can learn about the consequences of acid rain in the chapter on environmental chemistry.

10.4 CHARACTERISTIC PROPERTIES OF BASES

General properties of bases are as follows:

1. Bases have bitter taste.
2. Bases change the colour of red litmus paper to blue.
3. Aqueous solution of bases have slippery touch.
4. Bases neutralize acids to form salt and water.

$$H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$$

5. Bases decompose ammonium salts on heating and liberate ammonia gas.

$$\text{Base} + \text{Ammonium salt} \rightarrow \text{Salt} + \text{Ammonia} + \text{water}$$

$$NaOH + NH_4Cl \rightarrow NaCl + NH_3 + H_2O$$

$$KOH + NH_4Cl \rightarrow KCl + NH_3 + H_2O$$

10.5 OXIDES AND HYDROXIDES

Bases and hydroxides of metals consists of oxides and hydroxides of metals. Metallic oxides are the compounds formed by the reaction of metals with oxygen. For example, sodium reacts with oxygen to produce sodium oxide (Na_2O). Similarly magnesium on ignition in air burns producing magnesium oxide(MgO).

$$\text{Metal} + \text{Oxygen} \rightarrow \text{Metal oxide}$$

$$4Na + O_2 \rightarrow 2\,Na_2O$$

$$2\,Mg + O_2 \rightarrow 2\,MgO$$

When metal oxides dissolve in water, resulting in metal cations and oxide ions in aqueous solution. Because oxide ions are unstable in water, they immediately accept protons from water molecules and become hydroxide ions, and the water molecules also become hydroxide ions. Which species is the proton donor in this reaction?

$$O^{2-} + H\text{-}OH \rightarrow 2OH^{1-}$$

$$\textbf{Metal oxide + water} \rightarrow \textbf{Metal hydroxide}$$

$$MgO + H_2O \rightarrow Mg(OH)_2$$

$$Na_2O + H_2O \rightarrow 2\,NaOH$$

$$CaO + H_2O \rightarrow Ca(OH)_2$$

Most metal oxides and hydroxides are very basic in nature. They show the characteristic properties of bases. However, some metal oxides and hydroxides do not dissolve in water but behave in chemical reactions like both acids and bases. Such metal oxides and hydroxides are called amphoteric oxides and hydroxides, respectively. For example, aluminium oxide (Al_2O_3), alunimium hydroxjde ,$Al(OH)_3$, zinc oxide(ZnO) and zinc hydroxide($Zn(OH)_2$),etc.

KEY POINTS

- According to Arrhenius theory, an acid is a substance that ionizes in water to produce H+ ions and a base is a substance that ionizes in water to produce OH- ions.
- A Bronsted-Lowry acid is a proton donor and a base is a proton acceptor.
- A srong acid completely dissociates in aqueous solution and a weak acid dissociate partially in aqueous solution.
- A stong base completely dissociates in aqueous solution and a weak base dissociates partially in aqueous solution.
- Bases are oxides and hydroxides of metals.
- Alkalis are water soluble bases.
- Acids neutralize bases to form salt and water.
- Acids decompose carbonates to form salt, water, and carbon dioxide.
- Acids corrode metals and form salt and hydrogen
- References for additional information
- Longman Chemistry for IGCSE.
- IGCSE Chemistry.
- Cambridge IGCSE, Chemistry.
- Theories of Acids and Base Chemiguide.

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) Which of the following cannot be classified as Arrhenius acid?

(a)	HNO_3	(b)	H_2CO_3
(c)	CO_2	(d)	H_2SO_4

(ii) Which of the following is a Bronsted base?

(a)	NH_3	(b)	HCl
(c)	CH_3COOH	(d)	H_3O^+

(iii) Milk of magnesia contains $Mg(OH)_2$. It is used as antacid. It neutralizes excess stomach acid. Which salt is formed in this reaction?

(a)	$MgSO_4$	(b)	$MgCO_3$
(c)	$MgCl_2$	(d)	MgO

(iv) Ammonia is a base, because it

(a)	Ionizes in water to give OH^- ions	(b)	Contains OH group
(c)	Can accept an election pair	(d)	Can accept proton

(v) Consider the following reaction?

$$H_2O + HCl \rightleftharpoons H_3O^+ + Cl^-$$

(vii) Which species is an electron proton acceptor in this reaction?

(a) H_2O (b) HCl

(c) H_3O^+ (d) none

2. Give short answer.

(i) Write the equation for the self-ionization of water.

(ii) Define and give examples of Arrhenius acids.

(iii) Why HCl acts as a strong acid?

(iv) Why NH_3 acts as Bronsted-Lowry base?

(v) Why ammonia acts as a weak base.

3. Ammonium hydroxide and nitric acid react and produce ammonium nitrate and water. Write balanced chemical equation for this neutralization reaction.

4. Write balanced chemical equations for the following chemical reactions.

(i) Sulphuric acid + Magnesium hydroxide $\longrightarrow$ magnesium sulphate + water.

(ii) Sulphuric acid + Sodium Carbonate $\longrightarrow$ Soduimsulphhate + water.

(iii) Hydrochloric acid + Magnesium $\longrightarrow$ Magnesoum chloride + water

5. Identify Bronsted -Lowry acids or bases in the following reactions.

(i) $HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$

(ii) $NH_3 + HNO_3 \longrightarrow NH_4NO_3$

6. Give the Bronsted-Lowry definition of an acid. Write an equation that illustrates the definition.

7. Identify Bronsted acids and Bronsted bases in the following reactions.

Classify water as proton donor or proton acceptor.

(i) $CH_3COOH_{(aq)} + H_2O \rightleftharpoons CH_3COO^-{}_{(aq)} + H_3O^+{}_{(aq)}$

(ii) $HCO_3^-{}_{(g)} + H_2O_{(l)} \rightleftharpoons CO^{-2}_{3(aq)} + H_3O^+_{(aq)}$

(iii) $NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

(iv) $HCl_{(aq)} + HCO_3^-{}_{(aq)} \rightleftharpoons H_2CO_{3(aq)} + Cl^-{}_{(aq)}$

(v) $HS^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons S^{-2}{}_{(aq)} + H_3O^+{}_{(aq)}$

8. Write equations showing the ionization of the following as Arrhenius acids.
9. Write equations showing the ionization of the following as Arrhenius acids

(a) $HI_{(aq)}$ (b) $HNO_{2(aq)}$

THINK TANK

10. Compare the relative concentrations of hydrogen ions and hydroxide ions in each kind of solution?

(a) acidic (b) basic (c) neutral

11. Codeine, $C_{18}H_{21}NO_3$ is a commonly prescribed pain killer. It dissolves in water by the following reaction?

$$C_{18}H_{21}NO_3 + H_2O \rightleftharpoons [C_{18}H_{21}HNO_3]^+ + OH^-$$

Differentiate Codeine and water as Bronsted-Lowry acid or base.

12. Examine some ways in which you might determine whether a particular water solution contains an acid or a base.

PROJECT

Examine the labels of at least three antacid preparations. Make a list of the ingredients in each. Write a balanced chemical equation for the neutralization reaction that takes place when these antacids react with HCl in the stomach.

UNIT 11

ENVIRONMENTAL CHEMISTRY-AIR

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- State that composition of clean, dry air is approximately 78% nitrogen, 21% oxygen, remainder as a mixture of noble gases and carbon dioxide.
- State the major sources of air pollutants: Some examples include (a) carbon dioxide from the complete combustion of carbon-containing fuels (b) carbon mono-oxide and particulates from the incomplete combustion of carbon-containing fuels. (c) methane from the decomposition of vegetation and waste gases from digestion in animals (d) oxides of nitrogen from car engines (f) ground level ozone from reactions of oxides of nitrogen, from car engines, and volatile organic compounds, in presence of light.
- State the adverse effects of air pollutants: Some examples include (a) carbon dioxide; higher levels of carbon dioxide leading to increased global warming, which leads to climate change. (b) carbon mono-oxide as toxic gas (c) particulates: increased risk of respiratory problems and cancer (d) methane: higher levels of methane leading to increased global warming, which leads to climate change. (e) oxides of nitrogen: acid rain, photochemical smog, and respiratory problems (f) sulphur dioxide: acid rain and haze.

- Explain how the greenhouse gases carbon dioxide and methane cause global warming, some example include: (a) the absorption, reflection, and emission of thermal energy. (b) reducing thermal energy loss to space.
- Describe the role of sulphur in the formation of acid rain and its impact on the environment..
- Describe the strategies to reduce the effect of environmental issues. Some examples include (a) climate change: planting trees, reduction in livestock farming, decreasing use of fossil fuels, increasing use if hydrogen, and renewable energy, ,e.g., wind, solar (b) acid rain: use of catalytic converters in vehicles, reducing emissions of sulphur dioxide by using low sulphur fuels and flue gas desulphurization with calcium oxide.
- Describe the role of NO and NO_2 in the formation of acid rain, both directly and through their catalytic role in the oxidation of atmospheric sulphur dioxide.
- Explain how oxides of nitrogen form in car engines and describe their removal by catalytic converters, e.g. $CO + 2NO \longrightarrow 2CO + N_2$
- Define photosynthesis as the reaction between carbon dioxide and water to produce glucose and oxygen in the presences of chlorophyll and using energy from sun.
- Analyze how to use tools to reduce personal exposure to harmful pollutants (some examples include the usage of masks, air quality indices and CO detectors).
- Identify high risk situations in life including those where lomg-term exposure to these pollutants can lead to respiratory issues and reduction in quality and longevity of life.

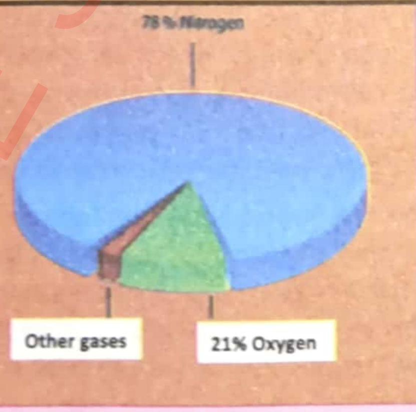
INTRODUCTION

This chapter will help you understand what atmosphere means. How the atmosphere is polluted and what substances pollute it? It is also necessary to understand the sources of general air pollution and the harmful effects on living things and the environment. You start your day with a cup of tea, which is a mixture of water, milk, tea, fat, and sugar. The same applies to milk, which is a mixture of water, minerals, proteins, vitamins, and fats. If scum or impure ingredients are added, the tea or milk is contaminated and unhealthy to drink. Air is a mixture of various gases including. nitrogen, oxygen, hydrogen, carbon dioxide, noble gases, and water vapor. The summer season is getting longer and hotter all over the world compared to ten years ago, why? You also know why iron nails rust faster in rainwater than in tap water or mineral water.

11.1 COMPOSITION OF ATMOSPHERE

Air is a mixture of gases. Can you name the main gases that make up the air? The pie chart given below shows the composition of dry air by volume. (Figure 11.1)

Composition of Dry Air



Other gases	% by volume
Argon	0.93
Carbon dioxide	0.038
Neon	0.0018
Helium	0.00052
Methane	0.00015
Krypton	0.00011
Hydrogen	0.00005

Figure 11.1 Composition of dry air

Besides gases, there are varying amounts of water vapour in the air. There is little water in the air over the desert. Whereas in the tropical rainforest, the air may contain up to 4% water vapour. This means the amount of water vapour in the air varies from place to place and from time to time.

The envelope of gases and water vapour surrounding the planet Earth is called the atmosphere.

CONCEPT ASSESSMENT EXERCISE 11.1

What two gases make up most of the air?

10.2 AIR POLLUTANTS

Think of a situation when you are in a park or a vegetable farm and in second case you are near a kiln or a garbage dump. Where would you feel fresh?

Pollutants are things like industrial wastes, herbicides, pesticides, insecticides, particles of dust and smoke, carbon monoxide, nitrogen dioxide, sulphur dioxide, ozone and lead containing paints. These things have a negative impact on the environment. Such substances effect environment as a result of human activity.

Anything that is in the air, water or soil which has a harmful effect on some part of the environment is called pollutant.

Pollutants damage the environment, health and quality of life. Important air pollutants are as follows:

11.2.1 Sulphur Oxides (SO_x)

You might have noticed that the colour of silk clothes fades away, if left in open air for a week or so. What due to it is?

Sulphur is found naturally in fossil fuels. The burning of fossil fuels in power plants, vehicles, industrial units, power generators and residential heating systems, therefore releases significant amounts of sulphur dioxide.

Sulphur dioxide is readily absorbed in the respiratory system. Being powerful irritant, it aggravates the symptoms of people who suffer from asthma, bronchitis, emphysema, and other lung diseases. Sulphur dioxide also responsible for acid rain and haze.

CONCEPT ASSESSMENT EXERCISE 11.2

1. What are pollutants?
2. List some effects of sulphur dioxide on human beings.
3. List some of the air pollutants.

The important oxides of nitrogen that cause air pollution are nitric oxide (NO) and nitrogen dioxide(NO_2). Collectively they are represented as NO_x.

The biggest source of nitrogen oxides is the burning of fossil fuels in vehicles, power plants, industrial units, and generators used to produce electricity. Car engines produce a lot of nitrogen oxides. Oxides of nitrogen highly toxic and are responsible for acid rain, photochemical smog, headache, and respiratory problems.

11.2.3 Carbon dioxide

Carbon dioxide (CO_2) is a major air pollutant that has received increasing attention due to its association with climate change and global warming. Although it occurs naturally in the Earth's atmosphere, human activities such as burning fossil fuels, deforestation, and industrial processes have dramatically increased its levels.

11.2.4 Carbon monoxide

Incomplete combustion, occurs when there is not enough oxygen in the combustion process. It is a major source of carbon monoxide and particulate matter. It is mainly emitted by vehicles, industrial processes and residential heating systems that use fossil fuels and wood. Particulate matter refers to a mixture of fine particles suspended in the air, including soot, smoke, dust and other solid or liquid substances. These particles can have harmful effects on human health, especially if they penetrate deep into the lungs, causing respiratory problems and lung cancer. Carbon monoxide is a toxic gas. When inhaled, it combines with hemoglobin to form carboxy hemoglobin, which is unable to carry oxygen causing you to lose consciousness and suffocate. It does not allow blood cells to absorb oxygen, and can cause death. Carbon monoxide is a colourless and odourless gas , so its presence cannot be felt.

11.2.5 Methane

Methane enters the atmosphere from a variety of sources, including the decomposition of vegetation and waste. Wetlands, rice fields, landfills, and livestock are important sources of methane emissions. Methane is also produced during the digestive process in animals such as cows, goats, and sheep. These animals have a unique digestive system that produces large amounts of methane during digestion, which is released through belching and flatulence. Higher levels of methane contribute to the global warming, which is responsible for climate change.

11.2.6 Ground Level Ozone

Ground-level ozone is often called smog. It is a secondary pollutant produced by complex reactions between nitrogen oxides and volatile organic compounds (VOC) under the influence of sunlight. These reactions occur mainly in industrial and urban areas where emissions from

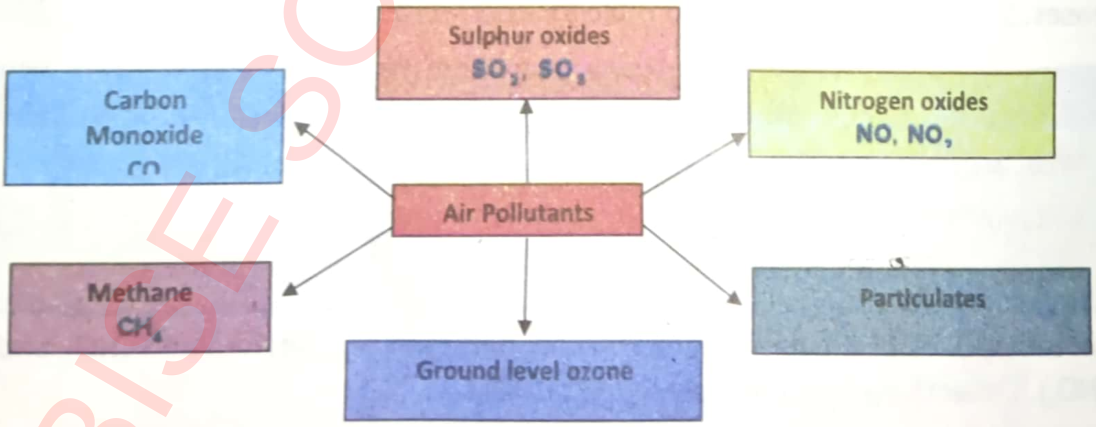


Figure 11.2: Air pollutants

industrial units and car engines are common. In sunlight, a series of chemical reactions occur between volatile organic compounds and nitrogen oxides. These reactions are called photochemical reactions. Sunlight breaks nitrogen dioxide into nitrogen monoxide and atomic oxygen. Atomic oxygen (O) then reacts with oxygen molecules (O_2) to form ozone (O_3). This ozone can irritate the respiratory organs and cause asthma and other respiratory problems. It can also damage vegetation, reduce yield and damage forests.

CONCEPT ASSESSMENT EXERCISE 11.3

1. Write the names of main pollutants in the air.
2. Complete the following reactions.

a) $SO_{2(g)} + O_{2(g)} \longrightarrow$

b) $C_{(s)} + O_{2(g)} \longrightarrow$

c) $CO_{(g)} + O_{2(g)} \longrightarrow$

11.3 SOURCES OF AIR POLLUTION

Air that contains harmful particles and gases is said to be polluted. Some air pollution occurs naturally. But many types of air pollution are the result of human activities.

11.3.1 Natural Sources

Many natural processes such as, forest fires, dust storms release smoke, and dust particles into the air. Volcano`s emit clouds of dust and poisonous gases along with ash. Which gas is emitted by volcanoes? Termites and cows also release large amount of methane in the air. Considerable electrical discharges in the atmosphere produce nitrogen oxides.

11.3.2 Human Activities

Most of the air pollution is the result of burning fossil fuels, such as coal, petroleum, and natural gas. Nearly half of the air pollution comes from cars and other motor vehicles. Factories and power plants that burn coal or oil release poisonous gases in the air. Burning fossil fuels and incineration release carbon monoxide (CO), nitrogen oxides (NO, NO_2), and sulphur oxides (SO_2, SO_3). Table 10.1 shows effect produced by air pollutants.

$$C_{(s)} + O_{2(g)}(\text{limited}) \longrightarrow CO_{(g)}$$

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

$$N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)}$$

$$2NO_{(g)} + O_{2(g)} \longrightarrow 2NO_{2(g)}$$

Table 11.1: Shows effect produced by air pollutants

Air pollutant	Physical properties	Sources	Harmful effects
Carbon monoxide	Colourless, odourless and poisonous gas	Incomplete burning of wood, fuels and vehicle exhaust.	Headache, brain damage, death.
Sulphur dioxide	Colourless gas with unpleasant and irritating odour	Power stations and industries using fossil fuels	Breathing difficulties, bronchitis, emphysema, lung cancer, acid rain and green house effect
Oxides of nitrogen	NO is colourless, odourless gas soluble in water. NO_2 is reddish brown gas with pungent odour soluble in water. Both are highly toxic gases	Exhaust fumes of motor vehicles, power stations and industries using fossil fuels	Coughs, headaches lung diseases, acid rain and greenhouse effect (global warming)

CONCEPT ASSESSMENT EXERCISE 11.4

1. Write three human activities that are responsible for air pollution.
2. Write three natural processes that are contributing in air pollution.
3. List main sources of the following air pollutants.
   (a) SO_2 (b) CO (c) NO_2

11.4 GLOBAL WARMING

The increasing use of fossil fuels and the deforestation have led to an increase in the levels of CO_2 in the air. Gases like water vapour, methane, and carbon dioxide act in a similar way in the atmosphere. These gases are called greenhouse gases.

When sunlight reaches the Earth's surface, some of it is absorbed by the Earth's surface, including vegetation, oceans, and land. This absorption warms the earth. As the earth's surface heats up, it emits energy in the form of infrared radiation. Greenhouse gases absorb this energy. These gases re-radiate some of the infrared radiation (see Figure 11.3).

DO YOU KNOW?

Green-houses are constructed from glass or transparent polymer films. Sun light can pass through these materials and is used by the plants for photosynthesis. The plants radiate some energy in the form of infrared or heat radiations which cannot pass through these materials and is reflected back. As a result the atmosphere inside the green-house becomes hot enough to promote plant growth. The temperature inside a greenhouse can be 10oC to 15oC higher than outside.

As a result greenhouse gases act like a blanket, trapping some of the heat energy in the lower atmosphere. This process is referred as greenhouse effect. This effect reduces thermal energy loss to space and helps to regulate the Earth`s temperature. Without this effect, the Earth would be much colder making it difficult for life to exist. Human activities have been increasing the concentration of greenhouse gases in the atmosphere, by burning fossil fuels, deforestation, and industrial processes. This additional increase in greenhouse gases is enhancing the natural greenhouse effect. This is leading to a significant increase in the atmospheric temperature causing global warming and climate change. The higher the concentration of greenhouse gases in the air, the greater the greenhouse effect and the greater the increase in temperature. The greenhouse effect is a natural phenomenon of the earth's energy distribution mechanism. **The warming of the atmosphere due to our influence on the greenhouse effect is called global warming.** Global warming is caused by disturbing the natural balance of greenhouse gas concentrations in the atmosphere.

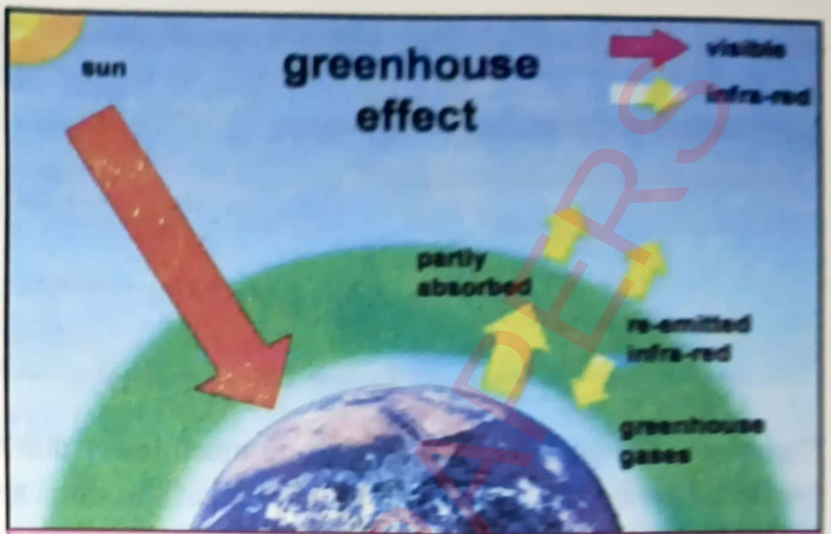


Figure 11.3 Greenhouse effect

If global warming continues, then

- Temperature of the earth will gradually increase.
- The earth climate may change, affecting both the rate of rainfall and how much it rains. This could cause both increased risks of flooding in some regions and drought in others.
- Polar ice may melt and cause significant increase in sea levels.
- So the atmosphere becomes hotter.

CONCEPT ASSESSMENT EXERCISE 11.5

1. Define global warming
2. List some effects of global warming
3. List some substances that are responsible for global warming.
4. Establish the link between greenhouse effect and global warming.

Society, Technology and Science

Incineration is a waste treatment process in which solid waste is burned at high temperature. Incineration consumes all combustible materials, leaving behind ash residue and non-combustible material. This process generally reduces the volume of waste by two third, but it is not a clean process. It produces air pollution. It generates considerable smoke and odour. This smoke may contain oxides of nitrogen and sulphur

11.5 ACID RAIN AND ITS EFFECTS

Normal rainwater is saturated with carbon dioxide. It has a pH of 5.6. However, the acidity of rain increases during thunderstorms in polluted areas. Fossil fuels contain compounds of sulphur an nitrogen. Sulphur dioxide from fossil fuels used in power plants and nitrogen oxides from ca exhausts dissolve in rainwater, creating acids, such as sulphuric acid, nitrous acid, and nitr acid. Nitrogen dioxide also play a significant role in the catalytic oxidation of sulphur dioxide into sulphur trioxide which forms sulphuric acid in the atmosphere. These reactions are key contributers to the acidity of rain water.

$$2SO_{2(g)} + O_{2(g)} \longrightarrow 2SO_{3(g)}$$

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$

$$2NO_{2(g)} + 3O_{2(g)} + 2H_2O_{(l)} \longrightarrow 4\,HNO_{3(aq)}$$

Therefore, the pH of rainwater can be much lower due to sulfuric acid and nitric acid from precipitation during thunderstorms. The pH of this rain can be as low as 2.1. This value is lower than the pH of vinegar or lemon juice.

Acid rain is defined as rain having pH less than 5.6.
Acid rain can often fall hundreds of kilometres from its source. Acid rain corrodes metals, stone structures and statues (Figure 11.4). Sulphuric acid eats metals and forms water-soluble salts and hydrogen.

$$Fe_{(s)} + H_2SO_{4(aq)} \longrightarrow FeSO_{4(aq)} + H_{2(g)}$$

Marble buildings and statues are disintegrated by acid rain.

$$CaCO_{3(s)} + H_2SO_{4(aq)} \longrightarrow Ca\,SO_{4(aq)} + H_2O_{(g)} + CO_{2(g)}$$

$$CaCO_{3(s)} + 2HNO_{3(aq)} \longrightarrow Ca\,(NO_3)_{2(aq)} + H_2O_{(g)} + CO_{2(g)}$$

Figure 11.4 Marble statues are slowly eroded by acid

Acid rain also kills fish, and destroys trees. Lakes and river may become too acidic for living things to survive. (Figure 11.5 and 11.6)

Figure 11.5: Trees destroyed by acid rain

Figure 11.6 Fish are killed by acid rain

CONCEPT ASSESSMENT EXERCISE 11.6

1. Define acid rain.
2. Write names of gases that cause acid rain.
3. What is the effect of acid rain on iron and marble? Give balanced chemical equation.
4. List some effects caused by acid rain.
5. Justify that acid rain is dangerous than normal rain

11.6 CATALYTIC CONVERTERS

Catalysts are installed in the exhaust system of cars. They contain catalysts that facilitate chemical reactions that convert nitrogen oxides into less harmful substances. The catalyst transforms CO to CO_2, NO to N_2 and O_2, and unburned hydrocarbons to CO_2 and H_2O. Metals such as platinum, palladium and rhodium are used as catalysts in the transformer.

$$2NO + 2HC/CO \rightarrow N_2 + 2CO_2 + H_2O$$

The government of Pakistan should direct car manufacturers to install catalytic converters in car exhaust systems. The government should make strict laws on this matter. This would help reduce emissions of air pollutants from car exhaust systems.

11.7 STRATEGIES TO REDUCE ENVIRONMENTAL PROBLEMS

How to address environmental issues? Several strategies that can be implemented are as follows:

1. Climate change migration

a) **Transition to renewable energy:** Increasing the use of renewable energy sources such as wind, solar, hydro and geothermal energy can help reduce greenhouse gas emissions.

b) **Energy efficiency:** promoting energy-efficient technologies can reduce energy consumption and related gas emissions.

c) **Reduce carbon dioxide emissions from transport:** Promoting electric vehicles and public transport systems and promoting walking and cycling can significantly reduce the use of fossil fuels.

2. Afforestation and reforestation.

a) **Planting trees:** Increasing plantations can help reduce carbon dioxide through photosynthesis. This practice not only improves air quality, but also prevents soil erosion. It also increases biodiversity.

b) **Forest Regeneration:** Look for initiatives to rehabilitate and restore damaged forests. This activity promotes carbon sequestration and wildlife habitat.

3. Sustainable land use.

a) **Sustainable agriculture:** Promotion of sustainable and renewable agricultural techniques, promotion of organic fertilizers instead of chemical fertilizers and pesticides. This practice can reduce emissions, protect water supplies and improve soil health.

b) **Reduce deforestation:** Taking steps to reduce deforestation and promote sustainable logging can help protect important carbon sinks.

4. Move to a low-carbon economy

a) **Transition to a hydrogen economy:** promoting the use of hydrogen as a clean energy source can help reduce carbon dioxide emissions in transport, industry, and power plants.

b) **Circular Economy:** Greenhouse gases can be controlled through measures to implement a circular economy model, minimizing waste and promoting recycling.

5. International cooperation and politics

a) **Paris Agreement:** To combat climate change, it is important to increase global cooperation and maintain the Paris Agreements, which include limiting global warming to well below 2°C and reaching the 1.5°C target.

b) **Policy support:** Implementation of strong and effective policies such as renewable energy incentives and emissions regulations can pave the way for visible change and promote sustainable practices.

c) It is important to note that these strategies are interrelated. They require a holistic approach involving individuals, businesses and public officials. However, cooperation, education and public awareness are important to achieve meaningful results.

11.7.2 Strategies to solve acid rain

Some approaches to dealing with acid rain include:

1. Acid rain mitigation:

a) **Emission control:** Measures to reduce emissions of sulfur dioxide, nitrogen oxides from industrial units, power plants, and vehicles are important for reducing acid rain. This can be achieved through the introduction of anti-pollution technologies and the use of cleaner fuels.

b) **Flue gas desulfurization:** Flue gas desulfurization systems, such as wet or dry scrubbers in power plants and industrial units, can remove sulphur dioxide from flue gases before they are released into the atmosphere.

c) **Low sulphur fuels:** Promoting low sulphur fuels can also minimize the amount of sulphuric acid in the atmosphere that causes acid rain.

2. Catalytic converters in vehicles:

a) The promotion of catalytic converters in vehicle exhaust systems can reduce emissions of harmful gases and volatile organic compounds. Converters contain catalysts such as platinum, palladium, and rhodium, which transform harmful gases into less harmful substances

11.8 PHOTOSYNTHESIS

Photosynthesis is a biochemical process in which plants use solar energy to convert carbon dioxide and water into glucose in the presence of chlorophyll. It is an essential process that sustains most life on Earth by producing oxygen and organic compounds. Photosynthesis plays a key role in the carbon cycle by removing carbon dioxide from the atmosphere and turning it into organic matter.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

11.9 PERSONAL PROTECTIVE MEASURES AGAINST POLLUTION

How to reduce personal exposure to harmful pollutants? Several tools and strategies can be employed. For instance:

1. Masks and Respiratory Protection

a) N95 masks are designed to filter out particles, including pollutants such as dust, smoke, and some airborne chemicals.

b) Gas masks contain activated carbon filters that absorb certain gases and volatile organic compounds.

NOTE. It is vital to choose masks certified by the relevant regulatory authorities and use precisely according to their instructions.

2. Air Quality Indices

Air quality indices provide information about the general air quality of a certain location. They usually use a numerical scale or a colour-coded system to indicate the level of air pollution. By regularly checking the air quality, you can take the necessary precautions.

3. Carbon monoxide detector

Installing carbon monoxide detectors in homes, offices and alerts people to carbon monoxide levels. This allows quick action to be taken, such as ventilating the area, sealing off potential sources or evacuating the area or enclosed spaces.

11.10 RISK FACTORS

High-risk life situations that can cause respiratory problems and reduced quality of life and life expectancy include:

1. **Long-term exposure to pollutants:** Breathing in polluted air containing nitrogen oxides, sulphur dioxide, particulate matter, ozone, and other harmful pollutants can damage the respiratory system. It can contribute to respiratory diseases such as asthma, lung cancer, and chronic obstructive pulmonary disease.
2. **Smoking:** Tobacco is another important risk factor for respiratory diseases. It damages the lungs and airways. This can increase the risk of chronic obstructive pulmonary disease and lung cancer. Cigarette smoke is also harmful to non-smokers.
3. **Indoor pollution:** Living in a home with poor ventilation, mold, dust, mites and pets, and household chemicals can cause respiratory problems, allergies and asthma.
4. **Allergens:** Prolonged exposure to allergens such as pollen, mold spores, pet dander, dust, and pet dander can cause respiratory allergies and asthma.
5. **Climate change:** Climate change affects air quality, which can worsen respiratory problems. Increased air pollution, forest fires, rising temperatures, and changes in pollen season can worsen respiratory symptoms. It is important to take protective measures such

as reducing exposure to pollutants, quitting smoking, maintaining good indoor air quality and wearing masks.

It is important to take protective measures such as reducing exposure to pollutants, quitting smoking, maintaining good indoor air quality and wearing masks.

KEY POINTS

- The envelope of gases and water vapour surrounding the planetearth is called atmosphere.
- Anything that is in the air, water or soil which has a harmful effect on some part of the environment is called pollutant.
- Some air pollution occurs naturally. But many types of air pollution are the result of human activities.
- Methane is produced when dead plant material decays in the absence of air.
- Air that contains harmful particles and gases is said to be polluted.
- The warming of the atmosphere which is due to our influence on the greenhouse effect is known as global warming.
- Acid rain is defined as rain having pH less than 5.6.
- Ozone is an allotropic form of oxygen comprising three oxygen atoms

References for additional information

- Chemistry in context.
- Chemistry, Kelter, Carr, Scott.
- Environmental Sciences, Cheris D.D. 1991.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which gas has the highest percentage in the air

(a) O_2 (b) CO_2

(c) N_2 (d) O_3

(ii) Which is/are responsible for acid rain?

(a) SO_2 (b) NO_2

(c) Both NO_2 and SO_2 (d) O_3

(iii) Which is reddish brown gas?

(a) NO (b) NO_2

(c) SO_2 (d) O_3

(iv) Most air pollution is caused by

(a) Ozone (b) Acid rain

(c) Carbon monoxide (d) The burning of fossil fuels

2. Give short answer.

(i) List two main sources of acid rain.

(ii) List four human activities which contribute to air pollution.

(iii) What is the importance of catalytic converters?

(iv) What is the role of automobile in air pollution?

(v) Define global warming.

3. Describe sources of air pollutants.

4. Describe acid rain and its effects.

5. Describe global warming.

6. What is ground level ozone. Explain.

7. Why is global warming often referred to as the greenhouse effect? Justify your answer.

8. Sulphur dioxide is a common pollutant from burning coal. State two effects caused by this pollutant.

THINK TANK

1. Dibenzothiophene ($C_{12}H_8S$) is a common sulphur containing compound of coal. It is responsible for acid rain. Elaborate this statement.
2. There is dire need to remove sulphur from coal before it is burned. Give reason.
3. Examine the option there are some ways to reduce pollution caused by cars?
4. Certain human activities are responsible for a significant increase in greenhouse effect, argue.
5. As a global citizen, how can you play a part to reduce air pollution at a personal level? Argue.

PROJECT

Global warming has become one of the most serious environmental issues in the world in recent times. Prepare a report on this issue in terms of:

(a) The gases contributing to the problem and their sources.

(b) Which of these gases are causing the most concern?

(c) Suggest some ways to reduce this problem.

UNIT 12

ENVIRONMENTAL CHEMISTRY-WATER

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Investigate chemical tests for the presence of water using anhydrous copper(II) sulphate.
- Explain how to test the purity of water using melting point and boiling point.
- Distinguish between distilled water and tap water with their applications in practical chemistry.
- State that water from natural sources may contain useful and harmful substances. (Some examples include:

  (a) Dissolved oxygen (b) Metal compounds (c) Plastics (d) Sewage (e) Harmful microbes (f) Nitrates from fertilizers (g) Phosphates from fertilizers and detergents)
- Recognize that some naturally occurring substances in water are beneficial. (some examples include:

  (a) Dissolved oxygen for aquatic life (b) Some metal compounds provide essential minerals for life)
- Recognize that some naturally occurring substances in water are potentially harmful (some example include:.

(a) Some metal compounds that are toxic (b) Some plastics that harm aquatic life (c) Sewage that contains harmful microbes which cause disease (d) Nitrates and phosphates that lead to de-oxygenation of water and damage to aquatic life. Details of the eutrophication process are not required)

- Explain the treatment of domestic water supply (some examples include: .
  (a) Sedimentation and filtration to remove solids (b) Use of carbon to remove tastes and colours (c) Chlorination to kill microbes
- Describe various water-borne diseases and the steps that can be taken to avoid them.
- Identify the negative effects of water pollutants on life and the ways to avoid them.
- Explain water scarcity as an important issue faced by Pakistan and the ways in which can be resolved.
- Fertilizers: State that urea, ammonium salts, and nitrates are used as fertilizers.
- Explain the use of NPK fertilizers to provide the elements nitrogen, phosphorous, and potassium for improved plant growth.

INTRODUCTION

Where does the water in your kitchen and bathroom come from? Because water comes from different sources, its quality varies. Do you drink water that has a colour, a bad taste or an unpleasant smell? When you wake up in the morning, what do you do? You brush your teeth, take a shower, flush the toilet, etc. What happens to the water flowing out of the drain? Can you reuse this water? Sewage is often poured into open gutters and allowed to flow directly into streams, rivers and oceans. This practice spreads diseases and also threatens aquatic life. How? This will become clear in this chapter. The presence of disease-causing bacteria affects water quality. Water from both public and private sources often requires treatment to ensure it is clean and safe to drink. Wastewater should also undergo treatment to remove unwanted substances before entering lakes, rivers or oceans. Otherwise, it would also affect marine life and, through the food chain, humans.

12.1 PROPERTIES OF WATER

- Water is the only substance that exists in three different states on Earth. Can you name these states?
- Pure water is transparent, colourless, odourless and tasteless. It boils at 100°C and freezes at 0°C at the sea level.

Activity 12.1

Detection of water with anhydrous(II) copper sulphate

Objective:

The objective of this activity is to demonstrate the ability of anhydrous copper(II) sulphate to detect the presence of water by observing its reaction and colour change.

Materials Required:

Anhydrous copper (II) sulphate (in powder form), a dry test tube or small beaker , a pipette or pipette, spatula.

Procedure:

Step 1:

Take a small amount of anhydrous copper(II) sulphate powder with spatula and place it in a dry test tube or beaker.

Step 2:

Observe the initial state and the appearance of anhydrous copper(II) sulphate powder before adding water. It should be a white or off-white crystalline.

Step 3:

Carefully add a few drops of water to the anhydrous copper(II) sulphate powder in the container using a dropper or pipette.

Step 4:

Observe the reaction. When water comes into contact with anhydrous copper(II) sulphate, a chemical reaction occurs in it. Copper(II) sulphate reacts with water to form hydrated copper(II) sulphate, which is blue in colour.

Step 5:

Observe and note the colour change that occurs. A white anhydrous copper (II)sulphate powder should turn light blue when reacted with water. This colour change confirms the presence of water.

Interpretation of results:

Anhydrous copper(II) sulphate is a white crystalline solid. It does not contain water molecules. However, it has ability to absorb water from the surrounding environment through a process called hydration. When water is added to anhydrous copper(II) sulphate, it undergoes chemical reaction. It reacts with water forming hydrated copper (II) sulphate, which has a distinct blue colour. The appearance of blue colour indicates the presence of water in it.

Safety Precautions:

Copper sulphate is a toxic compound, so handle it carefully. Perform this activity under the adult supervision. Wear gloves and safety goggles, and wash your hands, when the activity is over.

Activity 12.2

Testing the purity of water by determining its melting point.

Materials needed:

beaker, thermometer, bunsen burner or spirit lamp, ice cubes, tripod stand.

Procedure:

1. Make ice cubes of sample water using refrigerator.
2. Fill the glass with ice cubes.
3. Hang the thermometer on the ice cubes, making sure that its bulb is completely

immersed in ice and does not touch the bottom of the beaker.

4. Place the beaker on the heat source using the stand.
5. Gradually heat the ice cubes.
6. Record the temperature at which the ice begins to melt.
7. Record your observations.

Interpretation of results:

The melting point of pure water is 0°C. Impurities in water affect its melting point. Compare the observed melting point with the expected value.

Activity 12.3

Testing the purity of water by determining its boiling point.

Materials needed:

beaker, thermometer, bunsen burner or spirit lamp, water sample, tripod stand, a glass rod.

Procedure:

1. Fill the beaker with sample water and place it over the tripod stand.
2. Hang the thermometer in the water, making sure that its bulb is completely submerged in water and does not touch the bottom of the beaker.
3. Gradually heat water, constantly stirring water with a glass rod.
4. Record the temperature at which the water begins to boil.
5. Record your observations.

Interpretation of results:

The boiling point of pure water is 100°C. Impurities in water affect its boiling point. Compare the observed boiling point with the expected value.

12.1.1 Water as Solvent

Water is very good at dissolving substances. For this reason natural water such as rainwater and groundwater is not pure water. As water falls through the atmosphere, it dissolves, a little oxygen, nitrogen, carbon dioxide, and dust particles. During thunder storms, it also dissolves nitric acid. Ground water dissolves minerals from rocks and soils as it moves along on or beneath Earth's surface. Ground water also dissolves many substances from decaying plants and animals.

12.1.2 Quality of Water from Natural Resources

Water from natural resources such as lakes, streams, and underground rivers can contain many dissolved substances that can be beneficial or harmful.

12.1.3 Disadvantages of Natural Substances Found in Water:

1. Dissolved oxygen: Where do fish and other marine life get their oxygen? Water contains dissolved oxygen. This oxygen is respomsible for the survival of aquatic organisms.
2. Metal compounds: Natural water can contain metals such as iron, sodium, potassium, magnesium, calcium, manganese, zinc and copper. Some of these metals are necessary

for biological processes, their excessive concentration can be poisonous for aquatic organisms and humans.

3. **Plastics:** Plastic waste can pollute water bodies. This is a serious threat to aquatic life and ecosystem.
4. **Wastewater:** Wastewater can enter natural water supplies. Wastewater carries pathogens, bacteria, viruses and other harmful substances. These pollutants pose a serious threat to humans and aquatic life.
5. **Harmful Microbes:** Natural water sources can contain harmful microbes such as bacteria, viruses and parasites that can cause waterborne diseases such as diarrhea, dysentery, cholera and stomach upsets.
6. **Nitrates from fertilizers:** Fertilizers add nitrates to water bodies through agricultural runoff. High concentrations of nitrates in drinking water cause health risks, especially for young children.
7. **Phosphates from fertilizers and detergents:** Phosphates in fertilizers and detergents can enter water bodies through runoff. High concentrations of phosphate can cause eutrophication. This can cause harmful algae blooms and oxygen caps in water bodies. Detergents used in water systems can destroy the outer mucous membranes that protect fish from bacteria and parasites. In addition, detergents can also damage their gills.

12.1.4 Benefits of Natural Substances Found in Water:

Some naturally occurring substances are useful and necessary for life. For example,

1. **Dissolved oxygen:** Existance of aquatic life depends on dissolved oxygen. This oxygen supports the respiration process of aquatic organisms.
2. **Essential minerals:** Metal compounds in natural water can provide essential minerals needed for various biological processes. For example, iron is necessary for the production of hemoglobin in the blood, which carries oxygen to the cells of our body. Metals like zinc, copper, and manganese activate enzyme activity the proper functioning of biological systems.

12.2 TREATMENT OF DOMESTIC WATER SUPPLY

The treatment of domestic water supply involves several processes to ensure that water is safe for human use.

RAW WATER TREATMENT

Raw water is treated in a municipal water purification plant, to make it fit for drinking and domestic purposes. Various stages in this treatment are:

1. **Sedimentation:** It is the process in which water is allowed to stand in a reservoir. The suspended matter sinks to the bottom.
2. **Coagulation:** It is the process in which water is treated with slaked lime and alum. These materials react to form a gelatinous mass of aluminum hydroxide

$$3Ca(OH)_{2(aq)} + Al_2(SO_4)_{3(aq)} \longrightarrow 2Al(OH)_{3(s)} + 3CaSO_{4(aq)}$$

The aluminum hydroxide carries down dirt particles and bacteria.

1. **Filtration:** The water is then filtered through sand and gravel. Sometimes it is filtered through charcoal to remove coloured and odorous compounds.
2. **Chlorination:** In the final step, chlorine is added to kill any remaining bacteria. Chlorine reacts with water to form hypochlorous acid HClO which kills bacteria.

$$Cl_{2(aq)} + H_2O_{(l)} \longrightarrow HClO_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)}$$

Raw Water → Sedimentation → Coagulation → Filtration → Chlorination → Pure Water

Figure 12.1: Flow sheet diagram for water purification plant

12.2.1 Distilled Water

Water purified by distillation is called distilled water.

The process of distillation involves heating water to boiling point, collecting the vapors, and condensing it back into the liquid. The impurities are left behind. Distilled water is free of any contaminants. The electrical conductivity of distilled water is very low. The pH of distilled water is 7.

> **DO YOU KNOW?**
>
> Long-term drinking of distilled water can cause mineral deficiency because it lacks essential minerals found in natural water sources. For drinking purposes, it is better to use treated tap water, bottled water or natural spring water.

Uses in chemistry:

1. Distilled water is usually used in laboratories to prepare chemical reagents and solutions.
2. Distilled water is used in the calibration of sensors in analytical instruments, e.g. pH meter.

12.2.2 Tap water:

Water supplied through the municipal water system to households and for commercial purposes is called tap water. It comes from various sources, including, streams, rivers, lakes, reservoirs, and groundwater. It undergoes processes to meet safety standards to make it safe for people.

Applications in Chemistry Laboratory

1. Tap water is used as a solvent for general purposes, such as cleaning glassware, equipment, etc.
2. It is also used for educational and simple demonstration as a readily available solvent.

12.3 WATER POLLUTION

Water is very good at dissolving substances. As water from rain and snow flows over rocks and through soil, it dissolves minerals. The freshwater we drink or use for our daily life processes is a dilute solution containing a number of minerals. When these minerals are in sufficient concentration, water becomes unfit for human use. Many human activities also pollute the surface and groundwater. Human activities such as household waste, agricultural waste, livestock waste, pesticides, oil leaks, detergents, septic tanks, petroleum, and natural gas production may result in contamination of water bodies. We will discuss household waste and industrial waste in this unit. You will learn about other types of waste in higher grades.

Household Wastes

Household wastes include human wastes, livestock wastes, soaps and detergents, paints and oil, food, vegetable wastes, garbage, etc. Although detergents have stronger cleansing action than soap, they remain in the water for a long time and make water unfit for aquatic life. When household water containing detergents is discharged into lakes, ponds, rivers, etc. it causes the death of aquatic life. Chemical and bacterial contents in household water can contaminate surface and underground water. Bacterial contents present in water may cause infectious diseases such as cholera, jaundice, hepatitis, typhoid, dysentery, etc.

Society, Technology and Science

Water treatment is essential for many reasons.

- Through water purification, we can avoid drinking impure and containimated water, which causes many epidemic diseases and unsafe for healthy life.
- It removes bacteria, viruses and parasites which may cause diarrhoea, dysentery, botulism, typhoid, cholera, polio, and hepatitis.
- It also removes heavy metals like, As, Cr and Pb which can cause long term neurological problems, kidney diseases, nausea, dizziness, and cancer.
- It also improves the flavor and appearance.

12.4 WATERBORNE DISEASES

Human wastes are dumped on the ground or into the nearest stream. Human waste contains infectious microorganisms, which spread diseases like typhoid fever, dysentery, and hepatitis. Chemical and bacterial contents in livestock waste can pollute surface and groundwater causing the above-mentioned diseases. Hepatitis a viral disease occasionally spreads through drinking Water. Unclean water supplies, poor sanitation, and poor hygiene kill 2,668,000 people worldwide each year. Water in swimming pools is purified from pathogenic organisms by aeration and chlorination.

Some waterborne diseases are given below.

Dysentery

Dysentery is also an intestinal disease. It is caused by a parasite, entamoeba. This infection is transmitted by fecal contamination of water or food by the encysted organism. Patients have mild to severe abdominal cramps, diarrhea, chocolate-colored stool with mucous and sometimes blood.

DO YOU KNOW?

Chlorination is not effective against viruses such as those that cause hepatitis

Jaundice

This disease proceeds from obstruction of the liver. Excess of bile from the liver enters the blood and causes yellowness of skin and eyes. It leads to loss of appetite, weakness, and fatigue.

Hepatitis

Hepatitis is an acute inflammation of the liver. It is caused by viruses and is classified as Hepatitis A, B, C, D, and E. Hepatitis A and E spread through polluted water.

Society, Technology and Science

Swimming is an important recreational activity. Biological contamination has also lessened the recreational value of water. However, aeration and chlorination treatment of swimming pool water has lessened the threat of biological contamination.

Typhoid

Typhoid is a dangerous intestinal disease. It spreads by polluted water containing bacteria such as salmonella typhi, salmonella paratyphi, and salmonella enteritidis. It is characterized by continuous fever between 101 °F to 104 °F and irregular pulse.

CONCEPT ASSESSMENT EXERCISE 12.1

1. List some water borne diseases.
2. List sources of water borne diseases.
3. List steps used in raw water treatment

12.5 WAYS TO DEAL WITH THE NEGATIVE EFFECTS OF WATER POLLUTION

Water pollution affects life in many ways. Here are some ways to deal with them are as follows.

1. Some pollutants reduce oxygen levels in water bodies and make the survival of aquatic life difficult. To prevent this, it is important to control nutrient flow and properly treat sewage and industrial waste that encourage excessive algae growth.
2. Heavy metals present in polluted water, can accumulate in the tissues of aquatic organisms. These metals also harm human health if people consume them as food. Preventing industrial wastewater from entering waterways can minimize emissions of such pollutants.
3. Water-borne diseases like hepatitis, cholera, dysentery, etc. are caused by harmful microbes. Ensuring proper treatment of domestic and industrial wastewater, compliance with good sanitary and hygiene requirements, and providing access to clean drinking water are crucial in preventing water-borne diseases.

12.6 WATER SCARCITY IN PAKISTAN

Water scarcity is a major problem in Pakistan. This is mainly due to the following factors.

1. Population Growth:

The population of Pakistan is growing rapidly, which increases the demand for water in agriculture, industry, and homes.

2. Climate change:

Climate change has caused irregular rainfall patterns, which affects water availability. The country's water resources are also declining due to excessive extraction of water from groundwater aquifers and insufficient irrigation practices.

3. Inadequate water supply:

Traditional flood irrigation methods are widely used in Pakistan. This has resulted in significant water losses through evaporation and inadequate water distribution.

To solve the problem of water scarcity in Pakistan, the following are necessary:

a) Effective water management practices.

b) Development of infrastructure.

c) Policy reforms.

d) Public awareness of responsible water use and its promotion.

e) Strict regulations and monitoring systems to control groundwater and prevent illegal drilling are essential.

12.7 FERTILIZERS

Fertilizers are substances that provide essential elements for plant growth. These elements are essential to enhance crop yields. Fertilizers mainly provide three main nutrients, nitrogen (N), phosphorus (P), and potassium (K), hence known as NPK fertilizers. Urea, potassium nitrate, and ammonium salts such as di-ammonium phosphate (superphosphate) are important fertilizers. Fertilizers dissolve in water. So they provide nutrients to the plants in a readily available form. It is important to give fertilizers at the right time and in the right amount. This practice is called nutrient management. It optimizes the intake of nutrients by plants and reduces the loss of nutrients. This practice can minimize environmental problems such as nutrient runoff into water bodies.

KEY POINTS

- Hepatitis a viral disease occasionally spreads through drinking polluted water.
- Unclean water supplies, poor sanitation and poor hygiene kill 2,668,000 people worldwide each year.
- Anhydrous copper(II) sulphate is used to detect water; on absorbing water it turns blue.
- Water from natural resources contains useful and harmful substances.
- Nitrates and phosphates present in water damage aquatic life.
- Sewage contains harmful microbes which cause diseases.
- Sedementation, filteration, and chlorination are major steps in the treatment of domestic water supply.
- Urea, ammonium salts, and nitrates are used as fertilizers.

- NPK fertilizers provide the elements nitrogen, phosphorous, and potassium for plant growth.
- **References for additional information**
- Chemistry Kelter, Carr, Scott.
- Environmental Chemistry, Barid, Colin.
- Environmental Science, Richard Wright, R.T. W right.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which of the followings is not a water borne disease?

a) hepatitis b) typhoid
c) dysentery d) anemia

(ii) Which human activity results in contamination of water bodies?

a) livestock waste b) pesticides
c) septic tanks d) all of these

(iii) pH of distilled water is

a) 7 b) Less than 7
c) Greater than 7 d) None of these

(iv) Anhydrous copper (II) sulphate is ________ solid.

a) blue b) white
c) yellow d) None of these

2. Give short answer.

(i) List the impurities present in rain water.
(ii) List toxic substances present in household wastes.
(iii) In what ways, industrial wastes pollute water.
(iv) What is water pollution?
(v) List some waterborne diseases.
(vi) What are pathogenic microorganisms?

3. How does tap water differ from distilled water?
4. Why are municipal water supplies treated with aluminium sulphate?
5. What are some health effects of biological contamination of water?
6. Identify the negative effects of water pollutants on life.
7. Water scarcity is an important issue in Pakistan. Give your comments on it.
8. What are fertilizers?

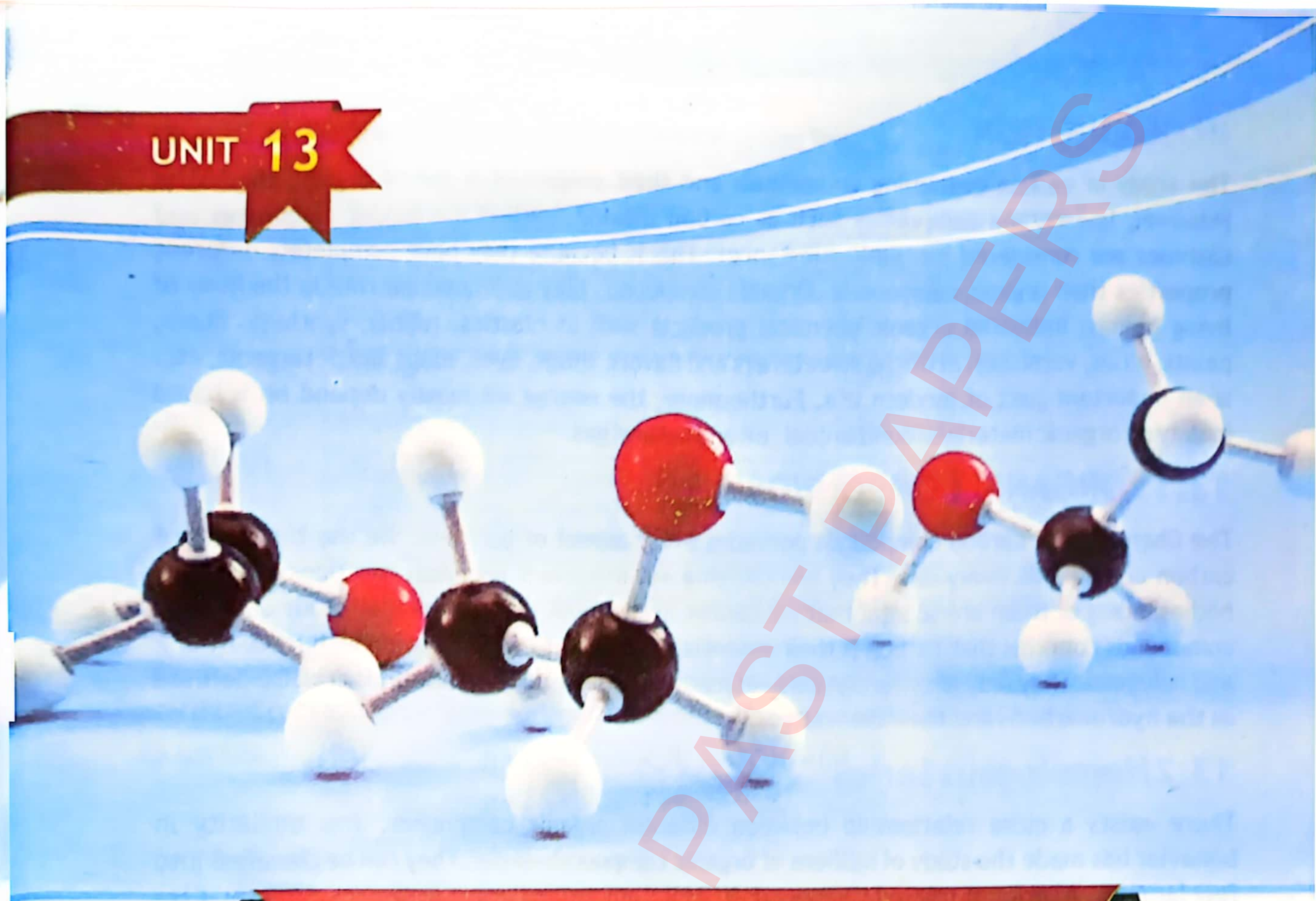
THINK TANK

9. Public health depends on water quality. Give arguments.
10. How chemistry helps maintain a clean swimming pool? Explain.
11. It is advisable to wash hands well with soap after using bathrooms. Evaluate it.

PROJECT

Create a chart showing water pollution model. (Watch YouTube - Kansal creation)

UNIT 13



ORGANIC CHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Describe organic molecules as either straight-chained, branched chained or cyclic.
- State that structural formula is an unambiguous description of the way the atoms in a molecule are arranged, including $CH_2=CH_2$, CH_3CH_2OH, CH_3COOCH_3.
- Identify and draw structural formulae for molecules.
- Interpret general formulae of compounds in the same homologous series including alkanes, alkenes, alkynes, alcohols and carboxylic acids.
- Define structural isomers as formula, but different structural formulae, including C_4H_{10} as $CH_3CH_2CH_2CH_3$ and $CH_3CH(CH_3)CH_3$ and C_4H_8 as $CH_3CH_2CH=CH_2$ and $CH_3CH=CHCH_3$
- Identify a functional group as an atom or group of atoms that determines the chemical properties of a homologous series including that for alcohols, aldehydes, ketones, phenols, carboxylic acids, amines, esters and amide.
- Describe the general characteristics of a homologous series. These can include: (a) having the same functional group (b) having the same general formula (c) differing from one member to the next by a $-CH_2-$ unit (d) displaying a trend in physical properties (e) sharing similar chemical properties.
- State that a saturated compound has molecules in which all carbon-carbon are single bonds.
- State that an unsaturated compound has molecules in which one or more carbon-carbon are not single bonds.

INTRODUCTION

The study of carbon-containing compounds and their properties is called organic chemistry. However, low carbon compounds such as carbon dioxide, carbon monoxide, carbonates and carbides are considered inorganic substances. This is because they have completely different properties than organic compounds. Organic compounds play an important role in the body of living beings. Industrial organic chemical products such as plastics, rubber, synthetic fibers, paints, glues, varnishes, artificial sweeteners and flavors, drugs, dyes, soaps and detergents, etc. is an important part of modern life. Furthermore, the energy we mostly depend on is based mainly on organic materials found in coal, oil and natural gas.

13.1 ORGANIC COMPOUNDS

The Chemistry of carbon compounds pervades every aspect of our lives. We use thousands of carbon compounds every day. They are carrying out important chemical reactions within our bodies. Many of them are so vital that we cannot live without them. A detailed study of organic compounds confirms that carbon is their essential constituent in combination with H, O, N, S, P and halogens. They may also (rarely) contain metal atoms. **Organic compounds are also defined as the hydrocarbons and their derivatives.**

13.2 Homologous Series

There exists a close relationship between different organic compounds. This similarity in behavior has made the study of millions of organic compounds easier. They can be classified into few families. A series of related compounds in which any two adjacent molecules differ by $-CH_2-$ group is called homologous series. For example, consider alkanes;

CH_4	CH_3-CH_3	$CH_3-CH_2-CH_3$	$CH_3-CH_2-CH_2-CH_3$
Methane	Ethane	Propane	Butane

Note the difference between adjacent alkanes, they differ by the same unit $-Ch_2-$ This means you can represent next member by simply adding $-CH_2-$ unit. A series of related compounds in which adjacent member differ by $-CH_2-$ is called a homologous series.

Similarly alcohols also form homologous series.

CH_3-OH	CH_3-CH_2-OH	$CH_3-CH_2-CH_2-OH$
Methanol	Ethanol	Propanol

These compound also differ by the same unit $-CH_2-$. All the classes of prganic compounds including alkanes, alkenes, alkynes, alcohols, aldehydes, ketones, carboxylic acids etc. form homologous series.

13.2.1 General Characteristics of a Homologous Series

Each homologous series have some general characteristics.

(i) The family members of a homologous series have the same functional group.

(ii) The family members have same general formula.

(iii) The adjacent family members differ by a $-CH_2-$ unit.

(iv) The family members display a trend in their physical properties.

(v) The family members possess similar chemical properties.

Table 13.1 shows the general formulae of some homologous series

Homologous series	General formula
Alkanes	C_nH_{2n+2}
Akenes	C_nH_{2n}
Alkynes	C_nH_{2n-2}
Alcohols	$C_nH_{2n+1}OH$
Carboxylic acids	$C_nH_{2n+1}COOH$

From the general formula you can easily determine the molecular formula of any member of the series. For examples,

Alkanes have general formula C_nH_{2n+2} which can be used to determine the molecular formula for any member of alkane series by putting number of carbon atoms in the general formula.

Examples: Methane $n = 1$ $C_1H_{2\times1+2} = C_1H_4 = CH_4$

Ethane $n = 2$ $C_2H_{2\times2+2} = C_2H_6$

Propane $n = 3$ $C_3H_{2\times3+2} = C_3H_8$

Akenes have general formula C_nH_{2n}

Ethene $n = 2$ $C_2H_{2\times2} = C_2H_4$

Alcohols have general formula $C_nH_{2n+1}OH$

Methanol $n = 1$ $C_1H_{2\times1+1}OH = CH_3OH$

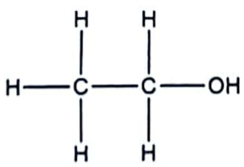
CONCEPT ASSESSMENT EXERCISE 13.1

Write the molecular formulae of the following compounds using general formulae

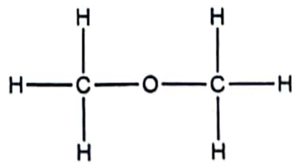
1. Alkane containing (i) 4 carbon atoms (ii) 6 carbon atoms
2. Alkene containing (i) 3 carbon atoms (ii) 4 carbon atoms
3. Alkyne containing (i) 3 carbon atoms (ii) 4 carbon atoms
4. Alcohol containing (i) 2 carbon atoms (ii) 3 carbon atoms

13.3 STRUCTURAL FORMULA

Frequently more than one organic compounds are represented by the same molecular formula. However, they have different properties. They have different structural formulas. For example, two organic compounds have the molecular formula C_2H_6O. They have different arrangements of atoms.

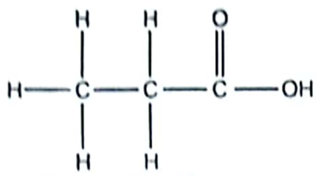


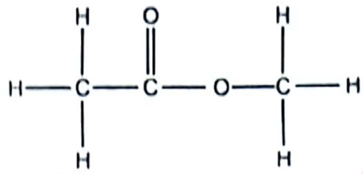
Ethanol



Dimethy ether

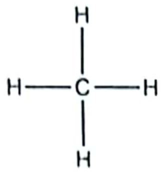
These formulas clearly show that the atoms are bonded to one another differently. **In ethanol,** the oxygen atom is bonded to only one carbon atom and a hydrogen atom. Whereas **in dimethyl** ether, the oxygen atom is bonded to two carbon atoms. Similarly, two organic compounds have the molecular formula $C_3H_6O_2$. They have different arrangement of atoms.

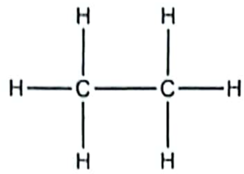

Propanoic acid

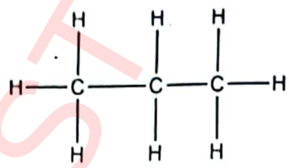

Methyl ethanoate

A formula that describes the arrangement of atoms in a molecule is called as structural formula.

The simple alkanes are straight-chain hydrocarbons. First three members of alkanes have following structural formulas.


Methane


Ethane


Propane

The condensed structural formulas of these alkanes are

$$CH_4\ ,\ CH_3CH_3\ ,\ CH_3CH_2CH_3$$

The corresponding molecular formulas are CH_4, C_2H_6, C_3H_8 respectively

A condensed formula is a structural formula that uses established abbreviation for various groups of chain. In condensed structural formula, we list the main chain carbon atoms and the hydrogen atoms attached to them in the sequence in which they appear in the naming system.

For instance,

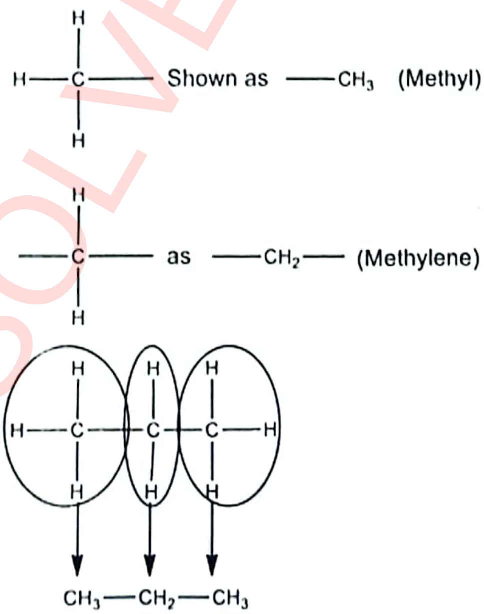


Table 13.2 shows the condensed structural formulas of some alkanes

Table 2.1: Properties of different states of matter

Name	Molecular Formula	Condensed Formula
Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$
Pentane	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$
Hexane	C_6H_{14}	$CH_3CH_2CH_2CH_2CH_2CH_3$
Heptane	C_7H_{16}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
Octane	C_8H_{18}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Nonane	C_9H_{20}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Decane	$C_{10}H_{22}$	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

Example 13.1: Give the molecular formula, the structural formula and the condensed structural formula for pentane

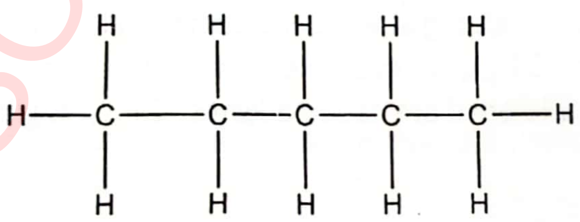
Problem Solving Strategy

1. The stem pent -means five carbon atoms.
2. The ending -ane indicates an alkane.
3. Write a string or chain of five carbon atoms.
4. Attach hydrogen atoms to the carbons to give each carbon atom four bonds. This requires three hydrogen atoms on each end carbon and two each on others.
5. For the condensed molecular formula, write each carbon atom's set of hydrogen atoms next to the carbon.
6. For molecular formula, simply count the carbon and hydrogen atoms or use the general formula C_nH_{2n-2} with n=5.

Solution:

C - C - C - C -C

Structural formula



Condensed Structural formula

$$CH_3—CH_2—CH_2—CH_2—CH_3$$

Molecular formula

$$C_5H_{2x5+2} = C_5H_{12}$$

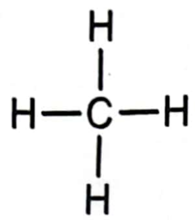
CONCEPT ASSESSMENT EXERCISE 13.2

Give the molecular, structural and condensed structural formulas for

(a) Butane (b) Hexane (c) Octane

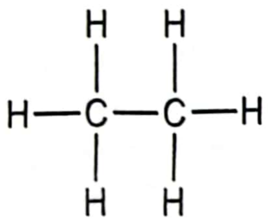
13.4 SATURATED AND UNSATURATED HYDROCARBONS

Hydrocarbons are compounds containing carbon and hydrogen only. **Hydrocarbons whose carbon - carbon bonds are all single bonds are called saturated.** Saturated hydrocarbons are also called alkanes. In alkanes each carbon atom is bonded to four other atoms. Methane is the simplest alkane. Other examples are ethane, propane, butane etc. The general formula of alkanes is C_nH_{2n+2}, where n is the number of carbon atoms.

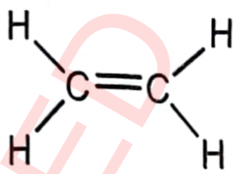


Methane

Hydrocarbons containing carbon-carbon multiple bonds are called unsaturated. Which of the following are unsaturated hydrocarbons



Ethane



Ethene

Ethyne

Unsaturated hydrocarbons are further divided into:

(i) Alkenes.

(ii) Alkynes.

Unsaturated hydrocarbons containing at least one carbon-carbon double bond are called alkenes. They have general formula (C_nH_{2n}), for example ethene. Unsaturated hydrocarbons that have at least one carbon-carbon triple bond are called alkynes. C_nH_{2n-2}, is general formula for alkynes, for example ethyne.

CONCEPT ASSESSMENT EXERCISE 13.3

Choose saturated and unsaturated compounds from the following.

(i) $CH_3-CH_2-CH_3$ (ii) $CH_3-C\equiv CH$

(iii) $CH_3-CH=CH_2$ (iv) $CH_2=CH-CH=CH_2$

13.5 CHEMICAL DIVERSITY OF ORGANIC COMPOUNDS

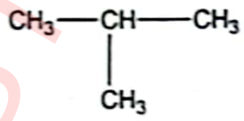
Carbon has four bonding electrons in its valence shell. Carbon, therefore forms four bonds with other atoms.

The Chemical diversity of organic compounds arises from carbon's ability to bond to each other to form long chains, branched chains and rings. This self-linking ability of carbon is called **catenation.** There appears to be almost no limit to the number of different structures that carbon can form. No other element can compete with carbon in this regard. Silicon and few other elements can form chains, but only short one. Carbon chains may contain thousands of carbon atoms.. Another reason for the large number of organic compounds is the phenomenon of isomerism.

The compounds that have same molecular formula but different arrangement of atoms in their molecules are called structural isomers. This phenomenon is called isomerism. For example two compounds have molecular formula C_4H_{10} but different structures.

$CH_3-CH_2-CH_2-CH_3$

n - Butane



iso - Butane

Alkenes also show structural isomerism. For example two alkenes have same molecular formula C_4H_8 but different structures.

$CH_3-CH_2-CH=CH_2$

1-Butene

$CH_3-CH=CH-CH_3$

2-Butene

13.6 CLASSIFICATION OF ORGANIC COMPOUNDS

There are millions of organic compounds. It is not possible to study each compound separately. To facilitate learning, they are divided into different groups and sub-groups. It is useful to choose those compounds with a similar structure. So here you will learn the classification of organic compounds based on the carbon skeleton. They are broadly divided into two main groups.

> **DO YOU KNOW?**
>
> Alkyl radical contains one less hydrogen than its parent alkane.

(i) Open chain compounds or Acyclic compounds.

(ii) Closed chain or Cyclic Compounds.

(i) Open chain compounds.

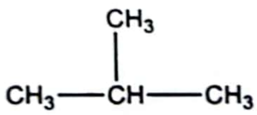
Open chain compounds contain an open chain of carbon atoms. In these compounds carbon atoms are linked in a linear pattern. For instance

$CH_3-CH_2-CH_3$

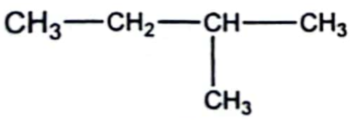
Propane

$CH_3-CH_2-CH_2-CH_3$

Butane

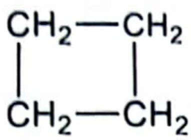


iso-Butane



iso - Pentane

Is the compound having following structure an open chain compound?



Open chain compounds may be either straight-chain or branched-chain. Those compounds which contain any number of carbon atoms joined one after the other in a chain or row are called **straight - chain compounds. In these compounds carbon atoms are connected in one continuous chain.**

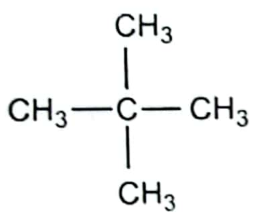
For example

$CH_3—CH_2—CH_3$ — **Propane**

$CH_3—CH_2—CH_2—CH_3$ — **Butane**

$CH_3—CH_2—CH_2—CH_2—CH_3$ — **Pentane**

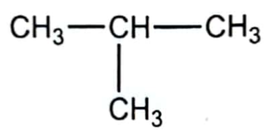
Those compounds which contain carbon atoms on the sides of chain are called branched chain compounds. In these compounds branches of carbon atoms are attached to a chain of carbon atoms. Which of the following is a branched chain compound?



neo-Pentane

$CH_3—CH_2—CH_2—CH_3$

Butane

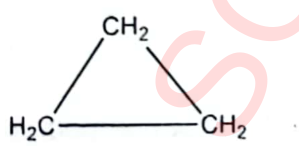


iso-Butane

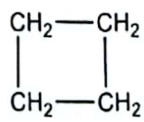
Open chain compound are also called alicyclic compounds

(i) Closed Chain or Cyclic Compounds

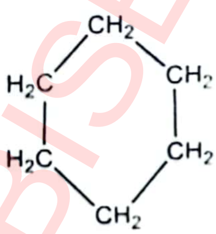
Organic compounds which contain rings of atoms are called closed chain or cyclic compounds. For example



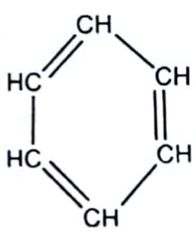
Cyclopropane



Cyclobutane



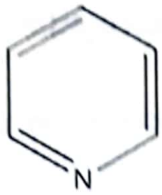
Cyclo hexane



or

Benzene

Cyclic compounds which contain rings of carbon atoms are called homocyclic or carbocyclic compounds. Which of the above cyclic compounds are carbocyclic? Cyclic compounds that contain one or more atoms other than carbon atoms in the ring are called heterocyclic compounds e.g.

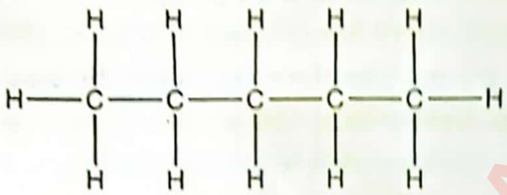

Pyridine

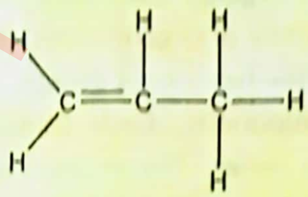

Furan

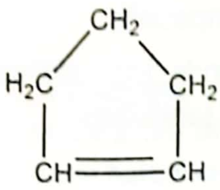

Thiophene

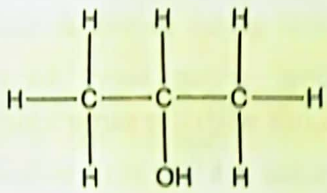
CONCEPT ASSESSMENT EXERCISE 13.4

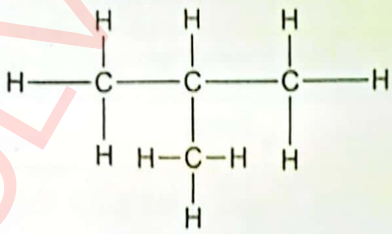
A to E are the structural formulas of some organic compounds.


A


B


C


D


E

Give the letters which represents

1. A branched chain compound.
2. A cyclic compound.
3. Two straight chain compounds.

13.7 FUNCTIONAL GROUPS

Most organic compounds contain elements other than carbon and hydrogen. Most of these compounds are considered hydrocarbon derivatives. That is, they are essentially hydrocarbons, but instead of one or more hydrogen atoms, they have an additional atom or groups of atoms called functional groups. In many simple molecules, the functional group is attached to an alkyl group.

An atom or groups of atoms that give a family of organic compounds its characteristic chemical and physical properties is called a functional group.

What is the difference in the following compounds?

CH_4	H_3C-OH	H_3C-Cl
Methane	Methyl alcohol	Methyl Chloride

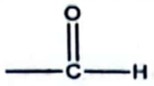
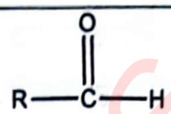
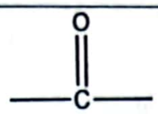
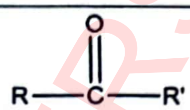
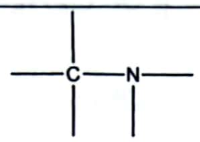
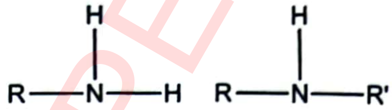
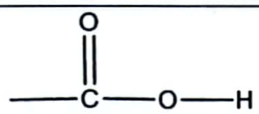
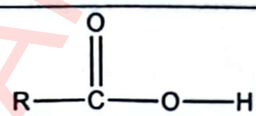
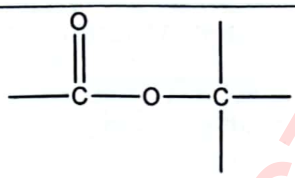
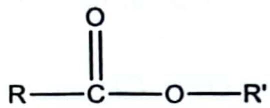
Research in organic chemistry is organized around functional groups. Each functional group defines a family of organic compounds. Although there are millions of organic compounds, there are only a few functional groups. Functional groups therefore facilitate the study of millions of organic compounds. Each functional group determines the characteristic properties of a homologous series. The properties of organic compounds are predominantly determined by the properties of the functional group present in a compound. The concept of functional group is important in organic chemistry for three reasons.

1. Functional group serves as basis for naming organic compounds.
2. Functional group serve to classify organic compounds into different classes. All compounds with the same functional group belong to the same class.
3. A functional is a site of chemical reactivity in a molecule.

The common functional groups are listed in the table 13.3

Table 13.3: Some Common Functional groups

Name of class	Functional group	General formula
Alkane	None	$R-H$
Alkane	$-\overset{\vert}{C}=\overset{\vert}{C}-$	$R-\overset{R'}{\overset{\vert}{C}}=\overset{R''}{\overset{\vert}{C}}-R'''$
Alkyne	$-C\equiv C-$	$R-C\equiv C-R'$
Alcohol	$-\underset{\vert}{\overset{\vert}{C}}-O-H$	$R-O-H$
Ether	$-\underset{\vert}{\overset{\vert}{C}}-O-\underset{\vert}{\overset{\vert}{C}}-$	$R-O-R'$

Aldehyde		
Ketone		
Amine		
Carboxylicacid		
Ester		

Each functional group exhibits characteristic properties

13.7.1 Functional groups containing Carbon, Hydrogen and Halogens: Haloalkanes

Haloalkanes are characterized by the presence of the halogen atom. The haloalkane is compound in which one hydrogen atom of an alkane is substituted by one halogen atom. Which of the following molecules are haloalkanes?

CH_4	H_3C-Cl	H_3C-Br
Methane	Chloromethane	Bromomethane

CH_3CH_2-I	CH_3CH_2-Cl
Iodoethane	Chloroethane

13.7.2 Functional groups containing Carbon, Hydrogen and Oxygen: Alcohols

Alcohols are characterized by the presence of the hydroxyl group. (-OH) attached to a hydrocarbon chain.

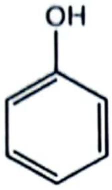
H_3C-OH	CH_3CH_2-OH
Methanol (Methyl alcohol)	Ethanol (Ethyl alcohol)

R - OH is the general formula for alcohols. Which of the following compounds is alcohol?

$CH_3CH_2CH_2OH$	$CH_3CH_2CH_2CH_2OH$
1-Propanol	1-Butanol

13.7.3 Phenols

When an - OH group is attached to a benzene ring, the compound is called a phenol.



or C_6H_5—OH

Phenol

Phenol was the first antiseptic used in an operation theatre.

13.7.4 Ethers

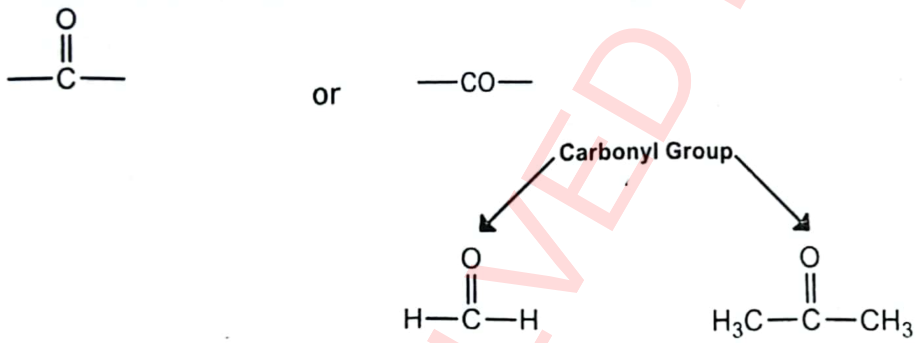
Organic compounds that have two alkyl groups attached to the same oxygen atom are called ethers. These compounds have C-O-C linkage in their molecules.

CH_3—O—CH_3	CH_3—O—CH_2—CH_3	CH_3—CH_2—O—CH_2—CH_3
Dimethyl ether	Ethyl methyl ether	Diethyl ether

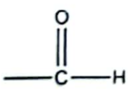
The general formula for ethers is R-O-R′. Where R and R′ are alkyl groups which may be same or different.

13.7.5 Aldehydes and ketones

Aldehydes and ketones contain the carbonyl group

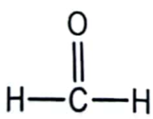


An aldehyde has at least one hydrogen atom attached to the carbonyl carbon atom. A ketone has two hydrocarbon groups (alkyl) bonded to the carbonyl carbon atom. Which of the above compound is an aldehyde? Which is a ketone?



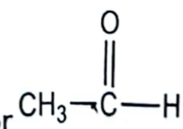
group in condensed form is written as -CHO. It is characteristic group of aldehydes

HCHO or



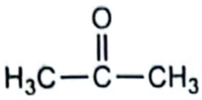
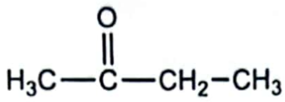
Methanal (Formaldehyde)

CH_3CHO or



Ethanal (Acetaldehyde)

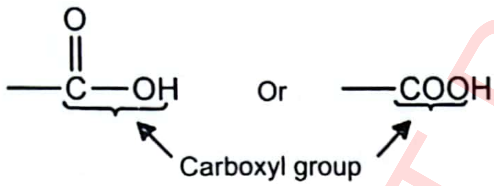
The general formula for ketone is $R-\overset{\overset{\displaystyle O}{\|}}{C}-R'$ and in condensed form it is written as $RCOR'$. Where R and R′ are alkyl groups which may be same or different. For example

 or CH_3COCH_3  or $CH_3COCH_2CH_3$

Propanone (Acetone) Butanone

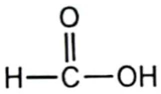
13.7.6 Carboxylic Acids:

The functional group of organic acid is called the carboxyl group.

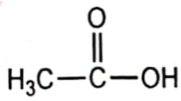


What is the difference between a carbonyl group and a carboxyl group ?

Examples:

 or HCOOH

(Formic acid) Methanoic acid

 or CH_3COOH

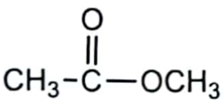
(Acetic acid) Ethanoic acid

The general formula for carboxylic acids is R - COOH

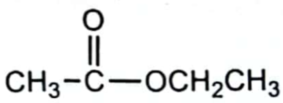
Where R = H or alkyl group

13.7.7 Esters:

Compounds having general formula $R-\overset{\overset{\displaystyle O}{\|}}{C}-OR'$ are called esters. R and R′ are alkyl groups which maybe same or different.



(Methyl acetate) Methyl ethanoate

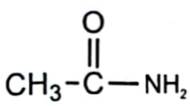


(Ethyl acetate) Ethyl ethanoate

$-\overset{\overset{\displaystyle O}{\|}}{C}-OR'$ is the functional group for esters.

13.7.8 Amides:

Compounds having general formula $R-\overset{\overset{\displaystyle O}{\|}}{C}-NH_2$ are called amides. R is an alkyl group.



ethanamide

13.7.9 Functional groups containing Carbon, Hydrogen and Nitrogen: Amines

The functional group of amines is $-NH_2$

CH_3-NH_2	$CH_3CH_2-NH_2$
Methyl amine	Ethyl amine

The general formula for amines is $R-NH_2$

Example 13.2: Differentiating different organic compounds on the basis of their functional groups.

Classify the following compounds as an alcohol, ether or a phenol.

1. $CH_3CH_2OCH_2CH_3$, is an anesthetic, but its' use as an anesthetic is now limited. This is because it is inflammable and causes nausea.
2. C_6H_5OH is a strong germicide. It is commonly used as disinfectant for floors, furniture and washrooms.
3. CH_3OH is poisonous and can cause blindness or death if taken internally.

Problem Solving Strategy:

1. Identify alkyl group in the molecule and functional group.
2. When -OH group is attached to an alkyl group, the compound is an alcohol, but when -OH is attached to benzene ring, the compound is a phenol.
3. When O- atom is attached to two alkyl groups, the compound is an ether.

Solution

(a) Ether (b) Phenol (c) Alcohol

CONCEPT ASSESSMENT EXERCISE 13.5

Classify the following compound as alcohol, ether or phenol.

(a) $CH_3CH_2OCH_2CH_3$ (b) $CH_3CH_2CH_2OH$

(c) C_6H_5OH (d) C_2H_5OH

Example 13.3: Classify the following organic compounds on the basis of functional group.

Identify the following compounds as an aldehyde or a ketone or a carboxylic acid.

1. CH_3COCH_3 is a common solvent for organic materials such as fats, rubbers, plastic and varnishes.
2. CH_3CH_2CHO has a foul irritating odour.
3. CH_3COOH is present in vinegar and is used to flavor food and making a polymer called polyvinyl acetate.

Problem Solving Strategy

Remember that

1. In an aldehyde a hydrogen atom is attached to the carbonyl carbon atom.
2. In a carboxylic acid -OH group is attached to the carbonyl carbon atom.
3. In a ketone, the carbonyl carbon is between two other carbon atoms.

Solution

(a) A ketone (b) An aldehyde (c) An organic acid

CONCEPT ASSESSMENT EXERCISE 13.6

Identify the following compounds as an aldehyde, or a ketone or a carboxylic acid.

(a) $CH_3COCH_2CH_3$ (b) $CH_3CH_2\overset{\overset{\large O}{\|}}{C}H$ (c) $CH_3CH_2\overset{\overset{\large O}{\|}}{C}OH$

Almost all synthesis involves the inter conversion of at least one functional group to another. A functional group is the active portion of the molecule. It plays a key role in the synthesis of new compounds. The key to design most organic synthesis is the functional group in the target molecules.

KEY POINTS

- The study of carbon-containing compounds and their properties is called organic chemistry.
- Organic Compounds are also defined as the hydrocarbons and their derivatives.
- A series of related compounds in which any two adjacent molecules differ by $-CH_2-$ group is called homologous series.
- A formula that describes the arrangement of atoms in a molecule is called structure formula.
- Hydrocarbons whose carbon-carbon bonds are all single bonds are called saturated.
- Hydrocarbons containing carbon-carbon multiple bonds are called unsaturated.
- The compounds that have same molecular formula but different arrangement of atoms in their molecules are called structural isomers.
- Open chain compounds contain an open chain of carbon atoms.
- Organic compounds which contain rings of carbon atoms are called cyclic compound.
- An atom or groups of atoms that give a family of organic compounds its characteristic properties is called functional group.
- **References for additional information**
- Chemistry for changing times, John W. Hill, Doris K. Kolb.
- Longman chemistry for IGCSE, Jin Clark and Ray Oliver.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Condensed structural formula for butane is

(a) $CH_3 - CH_2 - CH_3$ (b) $CH_3 - CH_2 - CH_2 - CH_3$

(c) $CH_3 - CH_2 - CH_2 - CH_2 - CH_3$ (d) $CH_3 - CH_3$

(ii) $CH_3 - CH_2 - CH_3$ is the chemical formula for

(a) Ethane (b) Propane

(c) Butane (d) Pentane

(iii) Which compound is not a saturated hydrocarbon?

(a) $CH_3—CH_3$ (b) CH_4

(c) $CH_3—CH{=}CH_2$ (d) $CH_3—CH_2—CH_3$

(iv) Stem "But" stands for how many Carbon atoms.

(a) 2 (b) 3

(c) 4 (d) 5

(v) The functional group $—\overset{\overset{\large O}{\|}}{C}—$ is found in

(a) Alcohols (b) Ketones

(c) Carboxylic acids (d) Esters

(vi) In which of the following Compounds, oxygen is attached to two alkyl carbon atoms?

(a) Alcohol (b) Phenol

(c) Ether (d) Ester

(vii) Which of the following is an alcohol?

(a) $CH_3 - CH_2 - O - CH_2 - CH_3$ (b) $CH_3 - CH_2 - COOH$

(c) $C_6H_5 - OH$ (d) $CH_3 - CH_2 - OH$

(viii) The functional group of amines is

(a) -OH (b) - COOH

(c) $-NH_2$ (d) - CHO

(ix) Ethanoic acid contains functional group

(a) - OH (b) - CO -

(c) -COOH (d) - CHO

2. Give short answer.

What is catenation?

(i) Define isomerism.

(ii) Give three examples of unsaturated compound.

(iii) Define a functional group.

(iv) What is the difference between an alkene and an alkyne?

3. Identify the following compounds on the basis of functional groups they contain and encircle the functional group.

$$CH_3-CH=CH_2,\quad CH_3-CH_2-COOH,\quad CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-CH_3,\quad CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-OCH_3,\quad H-\overset{\overset{\displaystyle O}{\|}}{C}-CH_3$$

4. What is the name of alkane having four carbon atoms in the chain?

5. Give the structural formula of two simple alkanes and one alkyne.

6. What is meant by the term functional group?

7. Identify the type of following compounds as an alcohol, aldehyde or ketone:

(a) HCHO, which is used to manufacture polymers, such as urotropine which is used to treat urinary tract infection.

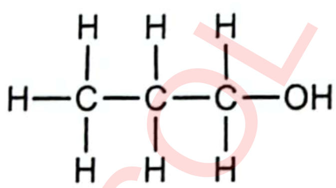
(b) CH_3COCH_3, which is used in nail polish remover.

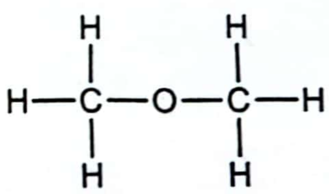
(c) CH_3CH_2OH, which is used in the preparation of many organic substances such as plastics, cosmetics, tinctures etc.

THINK TANK

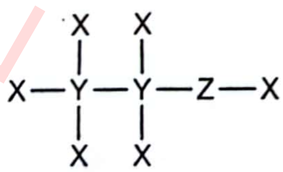
8. Give molecular formula of a compound containing C, H and O and single bonds. List all the possible functional groups this compound can have?

9. Give the condensed structural formulas of the following compounds and classify each on the basis of functional group.

(a) 

(b) 

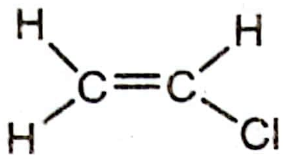
10. The diagram represents an organic compound that contains three different elements.



Select the possible compound from the following.

a) Ethanoic acid b) Propene c) Ehanol d) Propane.

11. Polyvinyl chloride (PVC) is a polymer. It is used for making vinyl sheets, drainage pipes, wire insulation etc. It is obtained from vinyl chloride

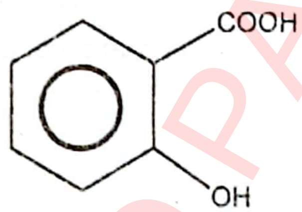


Classify Vinyl chloride as saturated or unsaturated compound.

12. For each of the following, sketch the structural formulas of a two-carbon compound containing the indicated functional group.

(a) alcohol (b) aldehyde (c)carboxylic acid (d) alkene

13. Aspirin is a mild pain killer and fever reducer. It is manufactured from salicylic acid.



Select functional groups present in it and encircle them. Justify your selection.

18. Construct the general formula for an alkane, an alkene, alkyne and an alcohol containing 4 carbon atoms.

19. Water adds to ethene according to the following reaction

$$CH_2{=}CH_2 + H_2O \longrightarrow CH_3CH_2OH$$

Compare the functional groups in the reactant and product molecules.

PROJECT

Prepare a chart showing differences between organic and inorganic compounds.

UNIT 14

HYDROCARBONS

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

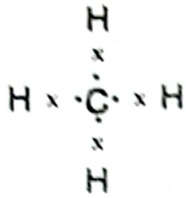
- State that bonding in alkanes is single covalent and that alkanes are saturated hydrocarbons.
- Describe the properties of alkanes as being generally unreactive, except in terms of combustion and substitution by chlorine.
- State that in a substitution reaction one atom or group of atoms is replaced by another atom or group of atoms.
- Describe the substitution reaction of alkanes with chlorine as a photochemical reaction, and draw the structural or displayed formulae of the products, limited to mono-substitution.
- Describe, using symbol equations, preparation of alkanes from cracking of larger hydrocarbons, hydrogenation of alkenes and alkynes, and reduction of alkyl halides

INTRODUCTION

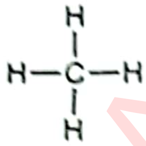
The simplest organic compounds are hydrocarbons. Organic compounds that contain only two elements, carbon and hydrogen, are called hydrocarbons. There are many types of hydrocarbons. They are classified according to the type of bond between the carbon atoms. In the previous chapter, you learned the differences between alkanes, alkenes, and alkynes.

14.1 ALKANES

Alkanes are saturated hydrocarbons. They have general formula C_nH_{2n+2}. Each carbon atom forms four bonds and each hydrogen atom forms only one bond. So the simplest alkane molecule that is possible is CH_4. It is called methane. Methane is the main component of natural gas. Electron dot and cross structure for methane is as follows.

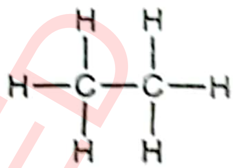


So the structural formula for methane is



Recall that structural formula shows which atoms are bonded to each other. The next member of alkane series is ethane, C_2H_6.

Ethane molecules has following structure



14.2 GENERAL METHODS OF PREPARATIONS OF ALKANES

1. By the hydrogenation of alkenes and alkynes

Addition of hydrogen molecule across carbon-carbon multiple-bond is called hydrogenation. Hydrogenation takes place in presence of finally divided nickel at 200 - 300°C and high pressure. Hydrogenation can also be done in presence of Pt or Pd at room temperature.

Alkynes add two molecules of hydrogen. Why?

$$CH \equiv CH + H_2 \xrightarrow{Ni, 200-300^\circ C} CH_2 = CH_2$$

Ethyne → Ethene

$$CH_2 = CH_2 + H_2 \xrightarrow{Ni, 200-300^\circ C} CH_3 - CH_3$$

Ethene → Ethane

CONCEPT ASSESSMENT EXERCISE 14.1

Complete the following reactions

$$CH_3 - CH = CH_2 + H_2 \xrightarrow{Ni, 200-300^\circ C}$$

$$CH_3 - C \equiv CH + 2H_2 \xrightarrow{Ni, 200-300^\circ C}$$

2. By the reduction of alkyl halides

When an alkyl halide is treated with Zn in presence of an aqueous acid, an alkane is produced. Usually aqueous solution of HCl or CH_3COOH is used. $CH_3-CH_2-Cl + 2[H] \xrightarrow{Zn/HCl} CH_3-CH_3$ Ethane + HCl

$$CH_3 - Cl + 2[H] \xrightarrow{Zn/HCl_{(aq)}} CH_4 + HCl$$

Zn reacts with aqueous acid to liberate atomic hydrogen called nascent hydrogen. Nascent hydrogen reduces alkyl halide. Addition of nascent hydrogen is called reduction. Long

CONCEPT ASSESSMENT EXERCISE 14.2

Complete the following reactions.

(a) $CH_3 - CH_2 - Cl + 2[H] \xrightarrow{Zn/HCl_{(aq)}}$

(b) $CH_3 - \underset{\underset{Cl}{|}}{CH} - CH_3 + 2[H] \xrightarrow[H_2SO_{4(aq)}]{Zn}$

3. By the cracking of larger hydrocarbons

A large hydrocarbon (alkane) molecule breaks into smaller hydrocarbons when heated at high temperatures such as 450-750°C and high pressure. This process is called thermal cracking. This process produces a mixture of alkanes and alkenes. For example, when decane is heated at high temperature and high pressure, it breaks down into octane and ethene.

$$\underset{\text{Decane}}{C_{10}H_{22}} \rightarrow \underset{\text{Octane}}{C_8H_{18}} + \underset{\text{Ethene}}{C_2H_4}$$

14.3 PROPERTIES OF ALKANES

Alkane molecules are essentially non-polar. They are less dense than water and do not dissolve in it. Chemically, alkanes do not react with most ionic compounds. The lack of reactivity makes them useful solvents. For example, hexane is used to extract vegetable oils from corn, soybeans, cotton seeds, etc. Alkanes containing up to four carbon atoms are colourless and odourless gases. Alkanes, containing five to seventeen atoms, are colourless and odourless liquids. Higher alkanes are solids that are also colourless and odourless. Long

1. Halogenation

Although unreactive towards ionic substances, alkanes readily react with halogens in sunlight. A chemical reaction that takes place in presence of sun light is called as photochemical reaction.

The reaction of an alkane and a halogen is a **substitution reaction.** In this reaction a halogen atom substitutes for one or more of the hydrogen atoms of an alkane.

For examples the reaction of methane and chlorine in diffused sunlight occurs as follows.

$$CH_{4(g)} + Cl_{2(g)} \xrightarrow{\text{diffused sunlight}} \underset{\text{Chloromethane}}{CH_3Cl_{(g)}} + HCl_{(g)}$$

In direct sunlight the reaction of methane with chlorine is explosive and forms carbon and HCl.

$$CH_{4(g)} + 2Cl_{2(g)} \xrightarrow{\text{direct sunlight}} C_{(s)} + 4HCl_{(g)}$$

2. Combustion

A reaction of a substance with oxygen or air that causes the rapid release of heat and the appearance of a flame is called combustion. Complete combustion of an alkane produces carbon dioxide, water and heat. Most of alkanes burn with blue flame.

For example, following reaction occurs when natural gas is burned.

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)} + \text{heat}$$

The lighter alkanes are widely used as fuel. This is because:

(i) Their combustion can be controlled.

(ii) They produce large amount of heat per gram.

(iii) They are cheap and readily available.

Incomplete combustion occurs in presence of limited supply of oxygen. Incomplete combustion of methane gives CO, C and H_2O.

$$3CH_{4(g)} + 4O_{2(g)} \longrightarrow 2CO_{(g)} + C_{(s)} + 6H_2O_{(g)}$$

KEY POINTS

- Hydrocarbons are compounds that contain carbon and hydrogen only.
- The simplest hydrocarbon that is possible is CH_4.
- Alkanes are generally unreactive
- Alkanes are saturated hydrocarbons
- In a substation reaction one atom or a group of atoms is replaced by another atom or a group of atoms.
- Addition of hydrogen molecule in an unsaturated hydrocarbon is called hydrogenation.

References for additional information

- Longman chemistry for IGCSE.
- Chemistry, Addison, Wesley. Fifth Edition.

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) Which molecule contains a carbon-carbon single bond?
(a)Ethane (b) Ethene
(c) Ethyne (d) Methanol

(ii) Which product is obtained when chloromethane (or methyl chloride) is reduced?
(a) Ethane (b) Ethene
(c) Methane (d) Ethyne

(iii) By hydrogenation we mean, the addition of
(a) Hydrogen (b) Water
(c) Halogen (d) Hydrogen halide

(iv) Combustion of methane produces
(a) Carbon dioxide (b) Water
(c) Heat (d) All of these

(v) Reduction of choloromethane gives
(a) Hydrogen (b) Cholorine
(c) Methane (d) All of these

2. **Give short answer.**

(i) Give three examples of saturated hydrocarbons.
(ii) Draw structure for ethane.
(iii) Draw structural formulas of an alkane containing five carbon atoms
(iv) What do you mean by hydrogenation reaction? Give one example
(v) What is meant by cracking?

3. Discuss methods for the preparation ethane.

4. Describe properties of alkanes.

5. Write a chemical equation to show the preparation of an alkane from an alkene and an alkyne.

THINK TANK

6. Write chemical equations for the preparation of propane

UNIT 15

BIOCHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Explain the importance and basics of nutrition and healthy eating.
- Recognize the main biomolecules, carbohydrates, proteins, lipids, and nucleic acids, their sources along with the required daily intake for young adults.
- Identify carbohydrates as a source of energy.

INTRODUCTION

Think about your brain. It works 24/7, even while you are asleep. Your brain supervises all your body functions, your thoughts, your breath, your heartbeat even your senses. In this scenario, your brain requires a constant supply of fuel. Food is a fuel that drives the processes of life. That "fuel" comes from the foods you eat and what's in that fuel makes all the difference. If your brain couldn't get good-quality nutrition, it can affect the functioning of brain. It means what you eat directly affects the structure and function of your brain and, finally, your life.

15.1 NUTRITION AND HEALTHY EATING

Nutrition is the source of food required by living things to stay alive. Our food choices impact our health. Good nutrition forms an important part of a healthy lifestyle. All living organisms are made up of molecules that carry out characteristics of life. The total number of molecules or nutrients that we need is called nutrition or diet.

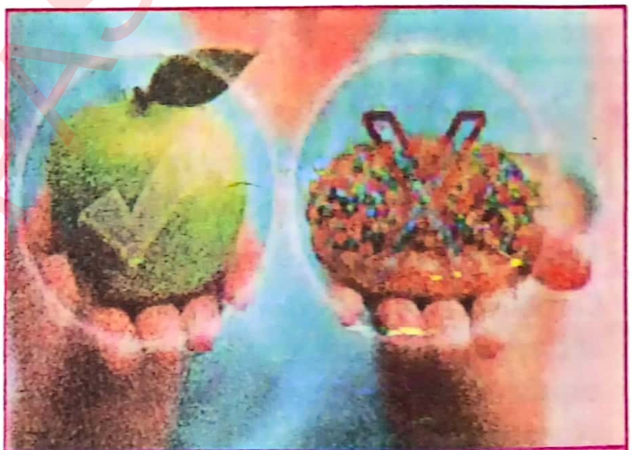


Figure 15.1: Healthy Diet

15.1.1 The Importance of Healthy Nutrition/Diet:

- Reduces high blood pressure.
- Lowers high cholesterol.
- A healthy balanced diet can prevent certain diseases, such as obesity, diabetes, cardiovascular diseases, cancer, and skeletal problems.

15.1.2 What is a Healthy Eating/Balanced Diet?

A "balanced diet" provides all the nutrients in suitable amounts needed to carry out the life processes. At every life stage, our diet is fundamental to our overall health and happiness. A healthy diet can lower the risk of disease. It also increases bone health and muscle strength.

A balanced diet includes a variety of foods in limited amounts and proportions to satisfy the needs for calories, proteins, minerals, vitamins, and other nutrients. A balanced diet should contain the correct proportion of six basic nutrients e.g., carbohydrates, lipids, proteins, vitamins, minerals, water, and dietary fibre.

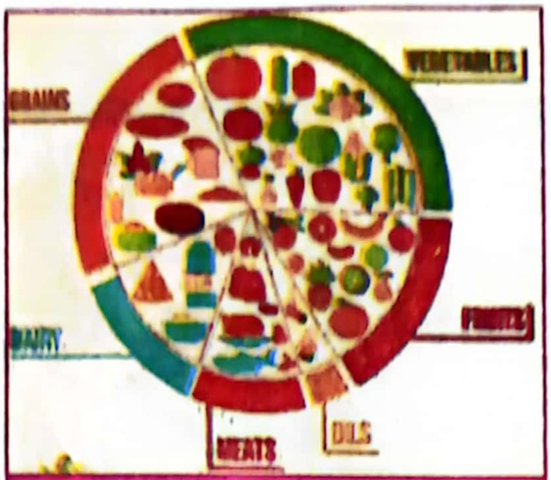


Figure 15.2: Healthy foods in daily routine

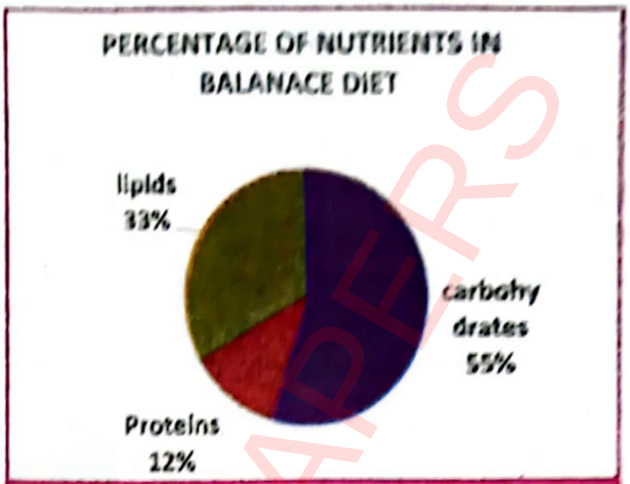


Figure 15.3: Percentage of nutrients in balance diet

15.2 BIOCHEMISTRY

Biochemistry is a combination of biology and chemistry. It is the study of living matter. The study of organic and inorganic molecules present in a living organism is known as biochemistry.

Carbohydrates, fats or lipids, protein and amino acid provide energy to the body that it needs to grow, play, and repair.

15.3 CARBOHYDRATES-A SOURCE OF ENERGY

Carbohydrates or hydrates of carbon are macromolecules made of carbon, hydrogen, and oxygen. The primary function of carbohydrates is to supply energy to all cells of body.

1 gram of carbohydrate $\longrightarrow$ 4 kilocalories energy.

Carbohydrates are broken down by the body into glucose - a type of sugar. Your body cells, tissues, and organs use glucose as fuel. When your body does not get suitable carbohydrates, it looks for another energy source, breaking down the protein in your muscles and body fat to use as energy. Glucose is the preferred fuel for the brain, which can't simply use other fuel resources like fat or protein for energy.

DO YOU KNOW?

A calorie is a unit that is used to measure the energy in your food. A calorie is the energy that it takes to raise the temperature of 1 gram of water by 1 degree centigrade. Calories are the amount of energy released when your body breaks down (digests and absorbs) food. The more calories a food has, the more energy it can provide to your body.

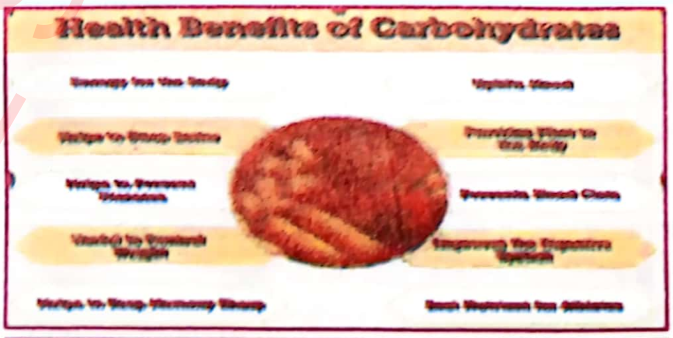
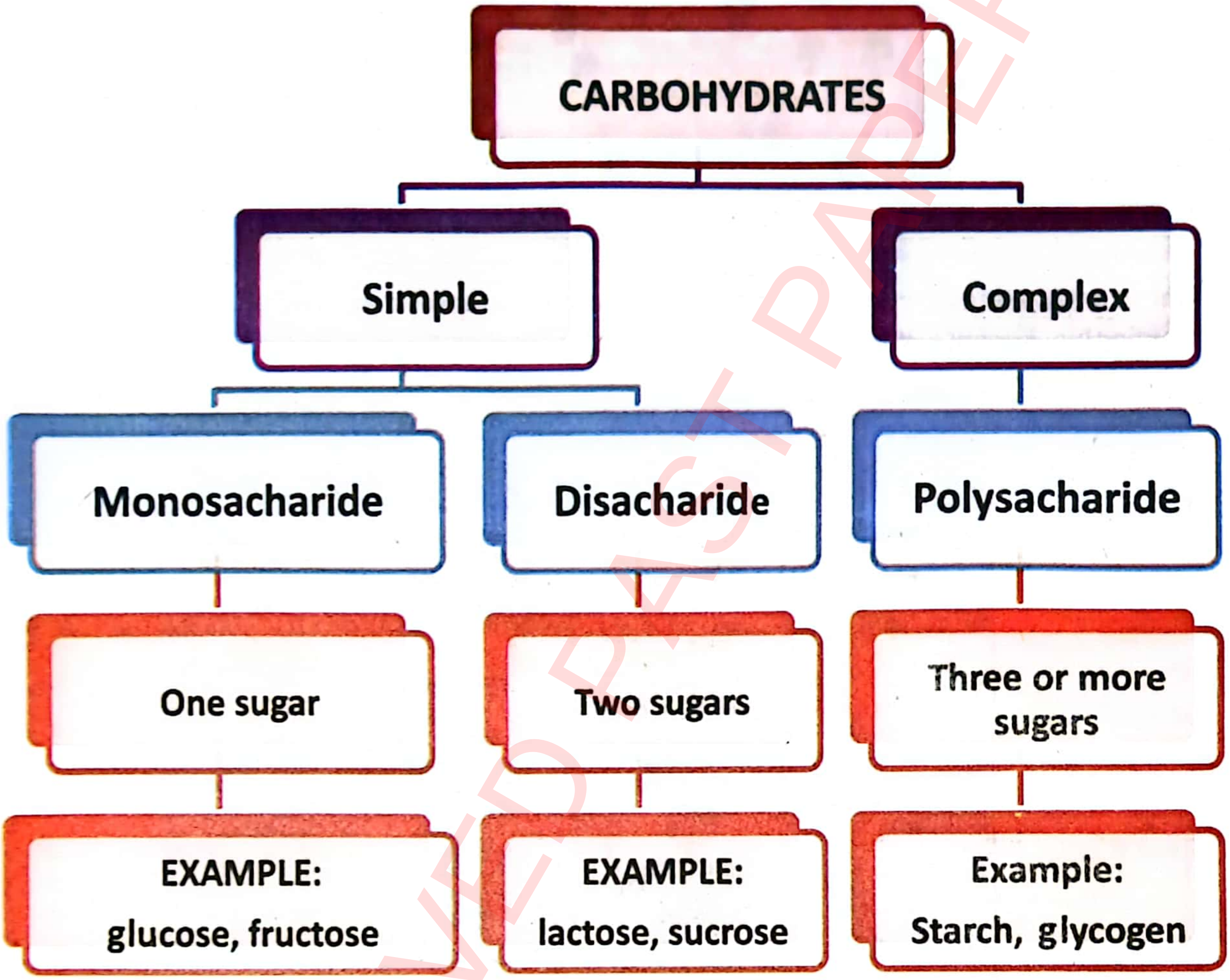


Figure 15.4: Benefits of carbohydrates

15.3.1 Classification of carbohyrates:

Carbohydrates can be classified based on how many sugars are in the molecule. The table below summarizes the three types of carbohydrates and examples.



15.4 PROTEINS:

Every cell in the human body contains protein. The basic structure of protein is a chain of amino acids. Proteins are macromolecules that are responsible for most of the functions that take place in the body. They do most of the work in cells and are required for the structure, function, and regulation of the body's tissues and organs. The building blocks of proteins are amino acids.

1 gram protein $\longrightarrow$ 4 kilocalories energy

DO YOU KNOW?

Enzymes are proteins that act as biological catalysts in living organisms. They help speed up chemical reactions in the human body.

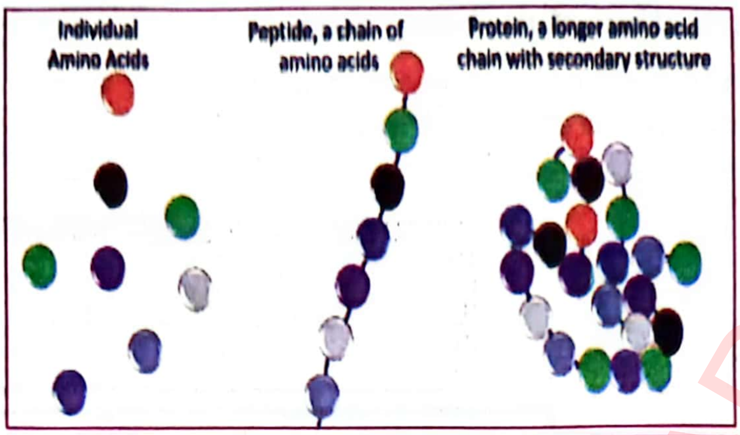


Figure 15.5: Structure of protein

Amino acids are molecules that combine to form proteins. Amino acids and proteins are the building blocks of life. When proteins are digested or broken down, amino acids are the result.

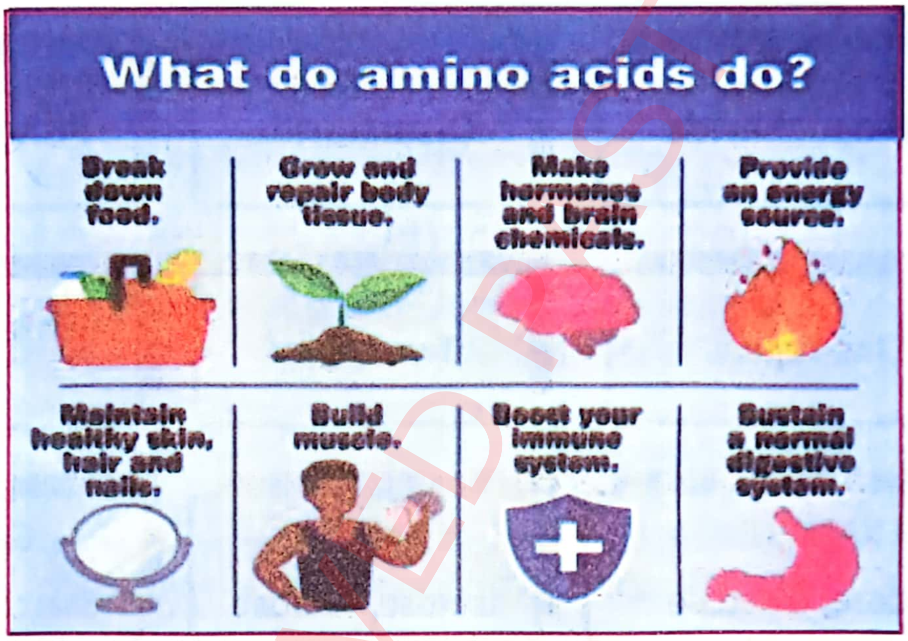


Figure 15.6: Functions of amino acids

15.5 LIPIDS:

Lipids are fatty, waxy, or oily compounds that are soluble in organic solvents like alcohol, chloroform and insoluble in polar solvents such as water. Lipids are mainly composed of hydrocarbons and little oxygen, making them an excellent form of energy storage. Lipids are fatty compounds that perform a variety of functions in your body. They're part of your cell membranes and help control what goes in and out of your cells. They help with moving and storing energy, absorbing vitamins, and making hormones.

DO YOU KNOW?

Where do lipids come from?
Most of the cholesterol in your body is produced by your liver. Your lifestyle choices, including diet, have a major influence on the production of cholesterol and triglycerides

Vitamin A, D, K, and E are the four fat-soluble vitamins and can be found in different foods like butternut squash, broccoli, and salmon fish. Eating dietary fat in a balanced diet helps you to absorb these fat-soluble vitamins, such as a cabbage salad with olive oil dressing and walnuts. Below are images of foods that contain the four fat-soluble vitamins.

1 gram of lipids provides 9 kilocalories energy

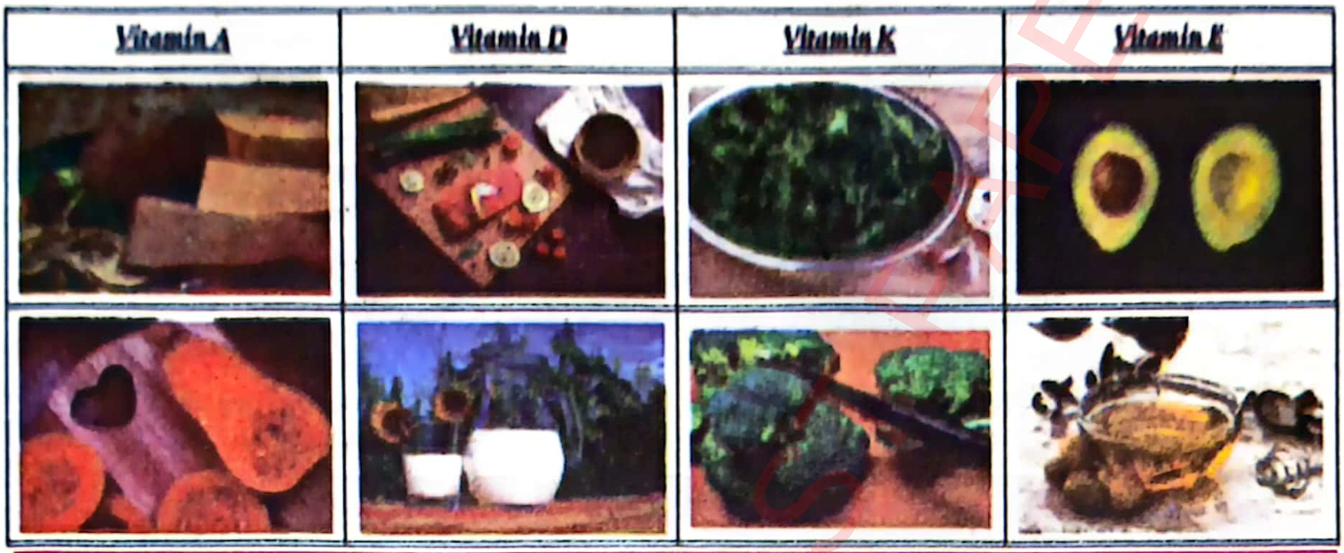
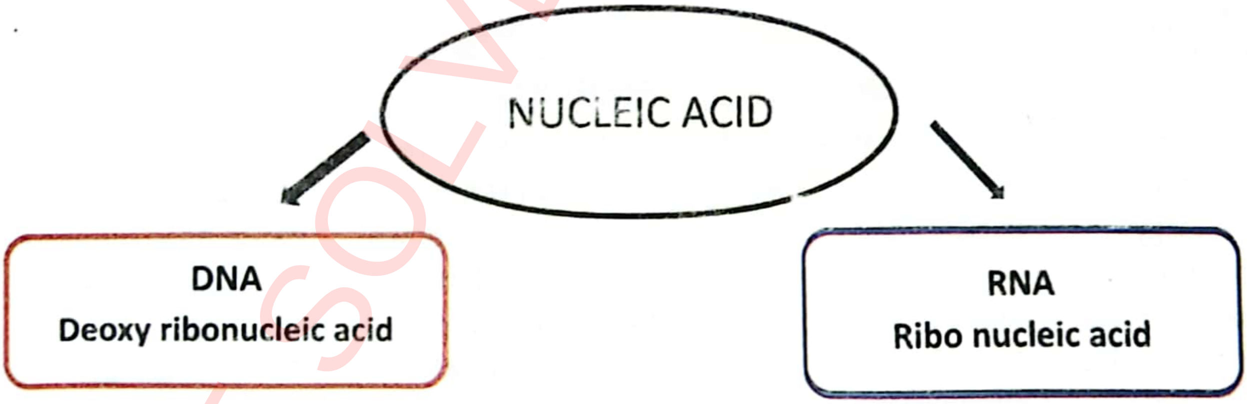


Figure 15.7: Foods that contain fat-soluble vitamins

15.6 NUCLEIC ACIDS:

Nucleic acids are large biomolecules (macromolecules) that play essential roles in the cells. Nucleic acids contain genetic information and play a key role in protein biosynthesis. They are macromolecules formed by nucleotides. Nucleotide is the basic building block of nucleic acids (RNA and DNA).



15.6.1 DNA:

DNA (deoxyribonucleic acid) is a double stranded form of nucleic acid that contains the genetic substance employed for the growth and functioning of all inferred living organisms

15.62.2 RNA:

RNA is a single-stranded nucleic acid that converts genetic information from genes into amino acid sequences of proteins.

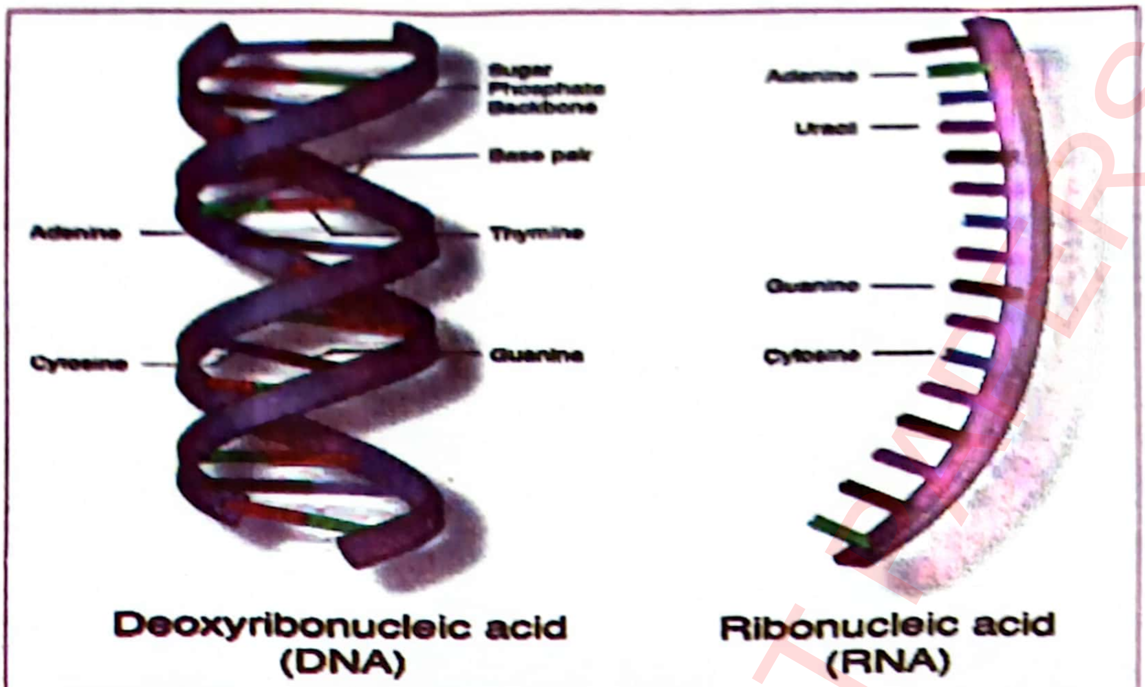


Figure 15.7: Nucleic acids

CONCEPT ASSESSMENT EXERCISE 15.1

1. Differentiate between DNA and RNA
2. Differentiate between proteins and lipids.

15.7 BASIC COMPONENTS OF FOOD:

CARBOHYDRATES			
SOURCE	**FUNCTION**	**PROPERTIES**	**DEFICIENCY**
Carbohydrates are the body's main energy source. Dairy products such as yogurt, milk, ice-cream, fruits, grains, plant-based proteins, and beans are major sources of carbohydrates.	Carbohydrates provide energy to feed the muscles and brain, while fiber helps to feel full and aids digestion and elimination.	Carbohydrates are soluble in water and readily broken down into sugars that can be oxidized during cellular respiration. These properties allow for carbohydrates to be a source of fuel for living things. Carbohydrates can be converted into fats and amino acids.	A carbohydrate-deficient diet may cause headaches, fatigue, weakness, difficulty concentrating, nausea, constipation, bad breath, vitamin and mineral deficiencies.

PROTEINS			
SOURCE	**FUNCTION**	**PROPERTIES**	**DEFICIENCY**
Proteins; are found in meat, fish, and eggsfrom animals and in peas, pulses, and beans from plants. Soya beans are one of the major sources. They contain very little fat.	The main role of proteins is to carry out functions for cells. Proteins are also essential for wound healing and for brain function. Proteins also help other types of communication within the body. Many hormones, or chemical messengers, are proteins. They speed up chemical reactions.	Proteins are colourless and usually tasteless. Proteins are digested in the stomach and small intestine and absorbed as amino acids. Proteins are used to make enzymes.	Little growth in children, brittle hair and nails, poor wound healing, anemia, unintentional weight loss, and scaly skin.
LIPIDS			
SOURCE	**FUNCTION**	**PROPERTIES**	**DEFICIENCY**
Meat and animal foods, e.g., eggs, milk, and cheese, are rich in saturated fats and cholesterol. Plant sources such as sunflower seeds and peanuts are rich in unsaturated fats.	Lipids primarily function as an energy reserve. Lipids are an important source of energy. They are insoluble in water. They also provide insulation to the body.	Lipids and fats are digested in small intestines and absorbed as fatty acids, and glycerol. The body can store unlimited amounts of fats contributing to obesity. lipids are not soluble in water but are thus soluble in solvents such as chloroform.	Dry rashes, hair loss, a weaker immune system, decreased growth in infants, and children, increased risk of different infections.

Nucleic acids			
SOURCES	**FUNCTION**	**PROPERTIES**	**DEFICIENCY**
Seeds, grain, and fish eggs are good sources of the genetic material, DNA.	DNA is a type of genetic material that contains all a person's genetic information. DNA also allows genetic information to be passed down from one generation to the next. RNA play a role in the expression of DNA's genetic code by generating certain proteins.	Nucleic acid carries genetic information which is read in cells to make the RNA and proteins by which living things work. Nucleic acids play a central role in a wide variety of cellular processes, including metabolic regulation and the storage and utilization of genetic information.	Lesch-Nyhan syndrome (a rare congenital (A birth disorder that affects a child's brain and behavior.), and ataxia telangiectasia also known as Louis-Bar syndrome, a rare inherited childhood neurological disorder that affects the part of the brain that controls motor movement, intended movement of muscles and speech.

15.8 REQUIRED DAILY INTAKE FOR YOUNG ADULTS:

A healthy diet requires lots of different nutrients in appropriate amounts that include fats, proteins, and carbohydrates; micronutrients such as vitamins and minerals; and an adequate amount of water to meet the needs for human nutrition.That means eating a variety of foods from each of the main food groups.

Consider the following tips for your meal:

•Eat more vegetables and fruit.

•Eat less carbohydrates.

•Eat less sugar.

•Eat less salt.

•Eat less fat.

The food pyramid shown here illustrates this diet.

Figure 15.8: Daily food intake for young adults

SUGGESTED HEALTHY EATING FOR TEENAGERS (AGED 12 TO 17)	SUGGESTED HEALTHY EATING FOOD FOR ADULTS
Grains: 4 - 6 bowls	Grains: 3 - 8 bowls
Vegetables: at least 3 servings	Vegetables: at least 3 servings
Fruits: at least 2 servings	Fruits: at least 2 servings
Meat, fish, egg, and alternatives: 200-300 grams	Meat, fish, egg and alternatives: 250 - 400 grams
Milk and alternatives: 2 servings	Milk and alternatives: 1- 2 servings
Fat/oil, salt, and sugar: Eat the least.	Fat/oil, salt, and sugar: Eat the least.
Fluid: 6 - 8 glasses	Fluid: 6 - 8 glasses

FACTS TO KNOW

Malnutrition is a condition that develops when the body is deprived of vitamins, minerals, and other nutrients. Malnutrition occurs in people who are either undernourished or over nourished.

Marasmus is a disease caused by insufficient levels of nutrition in the diet. Marasmus is severely undernutrition – a deficiency in all the macronutrients that the body requires to function, including carbohydrates, proteins, and fats

Research proves that food consisting of mostly high-calorie-density foods has a high risk of weight gain and obesity. Calorie density also affects hunger. Low-calorie-density foods tend to provide less fat and more water and fibre. Look at the density chart and examine:

1. Which food causes high blood pressure?
2. What are the dangers of high-calorie foods?
3. Why is it good to eat low-calorie foods? And why?

KEY POINTS

- A balanced diet describes all nutritional needs of the body.
- Carbohydrates, proteins, and lipids are biomolecules.
- Carbohydrates are the main energy source of the human diet.
- Lipids are organic compounds that are fatty acids or their derivatives. They are insoluble in water but soluble in organic solvents like chloroform. They include many natural oils, waxes, and steroids.
- Proteins are biomolecules that are responsible for most of the functions that take place in the body.
- Nucleic acids are large biomolecules that play essential roles in all cells.

REVIEW QUESTIONS

1. Encircle the correct answer.

(I) In small village children frequently suffered from different infectious diseases due to weak immune systems. Which major nutrients were lacking in their food?

(a) Amino acids (b) Carbohydrates
(c) Proteins (d) Lipids

(ii) Which of the following food components are rich in fats:

(a) Rice and maize (b) Pulses and wheat
(c) Milk, egg, and beans (d) Cheese, butter, and oil

(iii) Potatoes, cereals, beans, pulses, and oats are rich in:

(a) Proteins (b) Carbohydrates
(c) Amino acids (d) Fats

(iv) What is a bond between amino acids called?

(a) ionic bond (b) Acidic bond
(c) Peptide bond (d) Hydrogen bond

(v) The major function of carbohydrates includes:

(a) Storage (b) Structural framework
(c) Defense system of body (d) Messenger

(vi) Which of the following disorders is NOT caused by the deficiency of proteins?
(a) Weight loss (b) Muscle fatigue
(c) Loss in muscle strength (d) Constipation

(vii) Foods containing starch and carbohydrates are important because:
(a) They make your bones strong.
(b) They stop you from getting overweight.
(c) They are easy to cook. (d) They give you energy.

(viii) Meat, fish, and other alternatives provide the following important nutrients:
(a) Carbohydrates (b) Protein
(c) Lipids (d) Sugar

(ix) Which nutrient builds, maintains and repairs body tissues and cells?
(a) Carbohydrates (b) Protein
(c) Lipids (d) Water

(x) Marium's doctor told her that she is facing shortage of blood and have anemia. What nutrients are lacking in Marium's diet?
(a) Carbohydrates (b) Protein
(c) Lipids (d) Water

2. Give short answer.

(i) Give three reasons why living organisms need food.
(ii) What is a balanced diet? What is the importance of balance diet?
(iii) State four functions of proteins. Give one example to illustrate your answer.
(iv) Suggest two major foods a mother could give to her growing child? And why.
(v) Carbohydrates are a major source of energy. Defend the statement.
(vi) Fatima has fond of junk food like French fries, burgers, and pizza. What will happen if Fatima only eats junk food? What should she add to her routine food?

3. What are lipids? How are lipids important to our body?

4. How do you maintain a balanced diet?

5. What percentage of fat is required in a balanced diet? Why is the percentage of fat being lowest in major food components? Justify

6. What are the sources and functions of nucleic acids?

7. Imagine you are a nutritionist task to design a meal plan for athletes participating in a marathon. Explain the role of carbohydrates can affect an athlete's performance during the marathon.

PROJECT

Make a healthy meal diet plan and also identify food group of each food item.

UNIT 16

EMPIRICAL DATA COLLECTION AND ANALYSIS

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

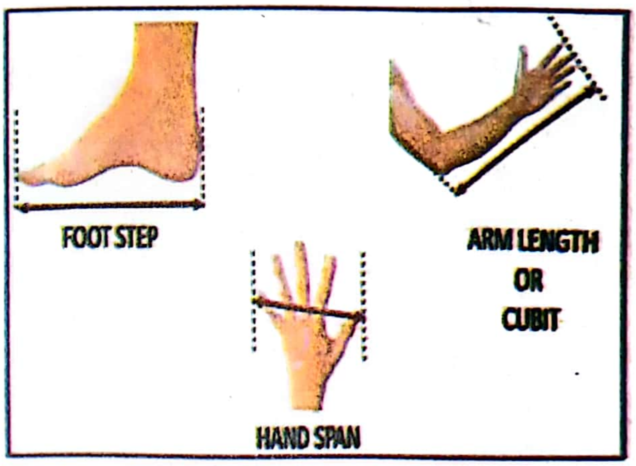
- Explain that units are standardized for better communication and collaboration
- Identify SI for abstract and physical quantities
- Apply the concept that units can be combined with terms for magnitudes, especially for kilo, Deci, and Milli.
- Justify why chemists use cm^3, g, and s as more practical units when working with small amounts in the lab.
- Explain with examples how different tools and techniques can be used to manage accuracy and precision for inherent errors that arise during measurement.

- Scientific narration standard form
- Use the standard form Ax10" where n is a positive or negative integer and is 1<A<10"
- Convert quantitative values into and out of the scientific notation form.
- Identify appropriate apparatus for measurement of time, temperature, mass, and volume, including
    1. Stopwatch
    2. Thermometer
    3. Balances
    4. Burrettes
    5. Volumetric pippetes
    6. Measuring cylinders
    7. Gas syringes
- Suggest advantages and disadvantages of experimental methods and apparatus

16.1 STANDARD UNITS

Measurement is an essential requirement to keep accuracy in our daily life. ***Measurement is a comparison of an unknown quantity with a known fixed quantity.*** For example, when you feel sick, your mother measures your body temperature with the help of a thermometer. The thermometer shows a 102°F temperature. So, do you have a fever? Yes, since our normal body temperature is 98.6°F, any value above it confirms that you are suffering from fever. It means that by measuring your body temperature, she will know exactly if you have a fever or not. So, Accurate, precise measurement is a fundamental component of chemistry.

In the past, people didn't have accurate measuring methods to calculate standard measurements. People used **abstract units** to measure with. For example, when an object was measured with a **cubit** or hand span, its length varied from person to person. The simple reason for the variation was the difference in the size of the cubit or hand-span of each person. Thus, this system of measurement was inconvenient as well as inaccurate. So, to maintain uniformity in measurement, scientists from all over the world accepted some of the units as standard units. This set of units is generally referred to as Standard International or SI system of units.



In 1960, scientists from different parts of the world gathered and agreed to adopt a single system of units called **the International System of Units** or **SI units.**

In the field of chemistry, the international system of units (SI) is used to measure physical quantities such as mass, volume, and temperature. This standard system ensures that chemists around the world can use the same units to measure and communicate their results, facilitating communication and collaboration in the field. Without standard units, it would be difficult for chemists to compare the results with one another, and it would be challenging to develop consistent and accurate scientific models.

During measurement, we compare the unknown quantity of things or substances with a known fixed quantity. This fixed quantity with which we compare unknown quantities is called a **unit of measurement**. The standard unit of measurement is a value that is fixed and cannot be changed.

Physical quantity:

A quantity that can be measured is called physical quantity. For instance, mass, amount of substance, length, time, temperature, etc.

16.1.1 SI UNITS

The entire metric system is composed of two different types of units, namely, fundamental units and derived units. The units are standard used to measure physical quantities, such as mass or length. There are seven units called Fundamental Units (or Base Units). These are used to measure different physical quantities. The kilogram, meter, and second are the fundamental base units. Whereas all other units that are derived from fundamental units are called **Derived Units**. So derived units are not independent; they are composed of two or more fundamental units.

Table 16.1: Fundamental units

S.no.	Physical Quantity	Unit	Symbol
1	Mass(m)	Kilogram	kg
2	Length (L)	Meter	m
3	Time (t)	Second	s
4	Temperature	Kelvin	K
5	Quantity of substance	Mole	mol
6	Electric current	Ampere	A
7	Luminous intensity	Candela	cd

16.1.2 SI Prefixes

The SI system develops a standard system of prefixes to the basic units, Prefixes are used to identify the original unit's multiples or fractions.

Table 16.2: Prefixes for measurements

Prefix	Unit abbreviation	Meaning	Example	Abbreviation
Kilo	k	1000	1 kilometer (km) = 1000 m	10^3m
hector	h	100	1 hectometer (hm) = 100 m	10^2m
Deca	da	10	1 decameter (dam) = 10m	10^1 m
			1 meter	10^0
Deci	d	1/10	1 decimeter (dm) = 0.1 m	10^{-1}m
Centi	c	1/100	1 centimeter (cm) = 0.01 m	10^{-2}m
Milli	m	1/1,000	1 millimeter (mm) = 0.001 m	10^{-3} m

Prefixes are also used to specify the number of atoms of each element in a molecule of the compound. When naming a binary molecular compound, the subscript for each element determines what prefix should be used.

The prefixes are written at the beginning of the name of each element. The following are the prefixes used for naming binary molecular compounds.

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

Binary Molecular Compounds

Formula	Name
CO	carbon monoxide
CO_2	carbon dioxide
SO_2	sulfur dioxide
NO	nitrogen monoxide
S_2Cl_2	disulfur dichloride
N_2O	dinitrogen monoxide
CCl_4	carbon tetrachloride
PCl_5	phosphorus pentachloride

16.2 CONVERSION AND THE IMPORTANCE OF UNITS:

The ability to convert from one unit to another is essential in scientific methods.

For example, a nurse has a tablet of 50 mg She has to give 0.2 g of tablet to a patient. She needs to know that 0.2 g equals 200 mg so 4 tablets are needed. There is a simple way to convert from one unit to another.:

LENGTH CONVERSION (Converting smaller units into larger units)					
10 millimeters	=	1 centimeter	10 mm	=	1 cm
10 centimeters	=	1 decimeter	10 cm	=	1 dm
100 centimeters	=	1 meter	100 cm	=	1 m
1000 meters	=	1 kilometer	1000 m	=	1 km

TIME CONVERSION (Converting smaller units into larger units)		
60 seconds	=	1 minute
60 minutes	=	1 hour
24 hours	=	1 day
7 days	=	1 week
365 days	=	1 year

MASS CONVERSION (Converting smaller units into larger units)		
100 milligrams	=	1 gram
1000 grams	=	1 kilogram

In the laboratory, chemists use cm^3 (cubic centimeters), g (grams), and s (seconds) as more practical units when working with small amounts in the lab for several reasons

- **Appropriateness:** The use of cm^3 allows chemists to easily measure and calculate the volume of liquids and solids. It is a smaller unit than liters and is more suitable for measuring small volumes accurately.
- **Precision:** To have precise measurements the use of grams as a unit of mass provides a more accurate and consistent measurement compared to larger units like kilograms.
- **Compatibility:** The use of these units ensures compatibility with a wide range of laboratory equipment and instruments. The use of smaller units makes it easier to perform experiments and obtain accurate results.
- **Time-sensitive reactions:** In chemistry, time plays a significant role in various reactions. Using seconds as a unit of time allows chemists to accurately measure reaction rates, reaction times, and other time-dependent parameters.
- **International Standards:** The International System of Units (SI) recommends the use of these units for scientific measurements. This standardization ensures uniformity and helps communication between scientists worldwide.

DO YOU KNOW?

The SI unit for volume is the cubic meter (m^3). The cubic centimeter (m^3), the liter (l), and the milliliter (ml) are also used.

1 liter(l) = 1000 milliliters (ml)

1 liter(l) = 1 decimeter cube (dm^3)

1 decimeter cube(dm^3) = 1000 cubic centimeters (cm^3)

1000 cubic centimeters (cm^3) = 1 liter (l)

16.3 SCIENTIFIC NOTATION:

Scientific notation or Standard form is used to represent very large or very small numbers in the form of multiplication of single-digit numbers and exponent raised to the power of 10. Scientific notation simplifies calculations, comparisons, and communication of measurement involving very large or very small numbers. It is a powerful tool for expressing the magnitudes of quantities used in everyday life.

For example, 650,000,000 can be written in scientific notation as 6.5×10^8.

The magnitude of any physical quantity can be expressed as:

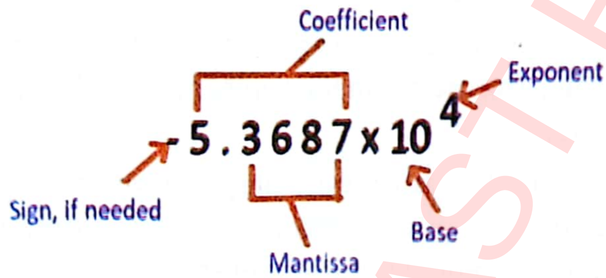
$a \times 10^n$

$1 \leq a < 10$ n is an integer

$A \times 10^n$

Where A is any number greater than or equal to 1 and less than 10 and n is any integer (whole number), negative or positive.

The exponent of base 10 determines how big or small the number. The exponent is positive if the number is very large, and it is negative if the number is very small. Numbers in **scientific notation or standard form** are expressed as a multiple of a power of ten.



Example 16.1:

Is the following number written in standard form?

$$14 \times 10^5$$

Solution:

A number is not written in standard form as A must be a number less than 10 and greater than or equal to 1. A is given as 14 which is greater than 10. This number is standard form would be:

$$1.4 \times 10^6$$

16.3.1 Scientific Notation Rules

To determine the power or exponent of 10, we must follow the rule listed below:

- The base should always be 10.
- The exponent must be a non-zero integer, which means it can be either positive or negative.
- The absolute value of the coefficient is greater than or equal to 1 but it should be less than 10.
- Coefficients can be positive or negative numbers including whole and decimal numbers.
- The mantissa carries the rest of the significant digits of the number.

Let us understand how many places we need to move the decimal point after the single-digit number with the help of the below representation.

- If the given number is greater than 10 then the decimal point has to move to the left, and the power of 10 will be positive.

  Example: $6000 = 6 \times 10^3$ is in scientific notation.

- If the given number is smaller than 1, then the decimal point has to move to the right, so the power of 10 will be negative.

  Example: $0.006 = 6 \times 0.001 = 6 \times 10^{-3}$ is in scientific notation.

Scientific Notation Examples

$$490000000 = 4.9\times10^{8}$$
$$1230000000 = 1.23\times10^{9}$$
$$50500000 = 5.05 \times 10^{7}$$
$$0.000000097 = 9.7 \times 10^{-8}$$
$$0.0000212 = 2.12 \times 10^{-5}$$

Example 16.2:

Convert 0.00000046 into scientific notation.

- **Solution:** Move the decimal point to the right of 0.00000046 up to 7 places.
- The decimal point was moved 7 places to the right to form the number 4.6.
- Since the numbers are less than 1d the decimal is moved to the right. Hence, we use a negative exponent here.

$$\Rightarrow 0.00000046 = 4.6 \times 10^{-7}$$

This is the scientific notation.

EXAMPLE 16.3: Convert 12 kilograms to milligrams.

Solution:

To convert from a larger unit to a smaller unit, we need to multiply.

There are 1000 grams in 1 kilogram.

Therefore, 12 kilograms will have

12 ×× 1,000 = 12,000 grams

So there are 12,000 grams in 12 kilograms

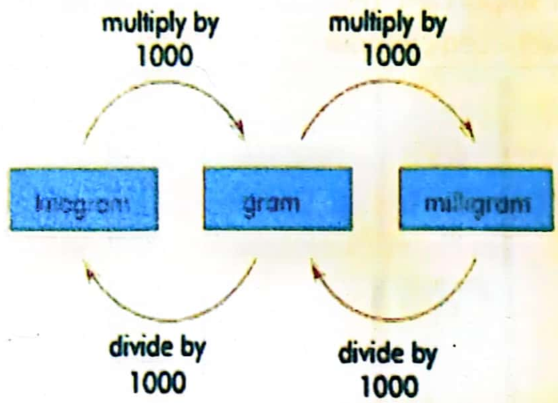
Now, there are 1000 milligrams in 1 gram.

Therefore 12 kilograms will have,

12,000 ×× 1,000 = 12,000,000 milligrams

So, there are 12,000,000 milligrams in 12,000 grams.

Hence, there are 12,000,000 milligrams in 12 kilograms.



16.3.2 Converting Ordinary Form to Standard Form:

EXAMPLE	ORDINARY NOTATION	STANDARD NOTATION
Diameter of Earth	12 700 000 m	1.27×10^{7} m
Length of the great wall of china	6 400 000 m	6.4×10^{6} m
Height of a soldier	1.7 m	1.7×10^{0} m
Length of mosquito	0.01 m	1×10^{-2} m
Length of red blood cell	0.0 000 075 m	7.5×10^{-6} m

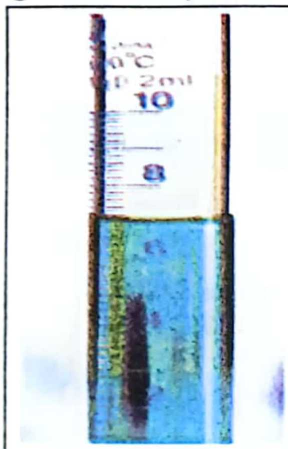
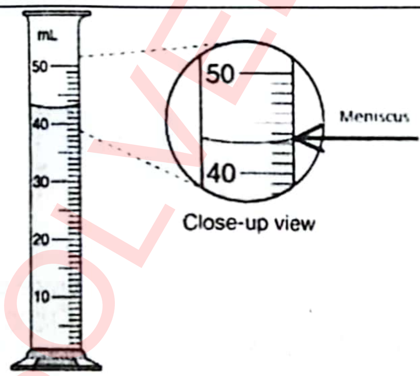
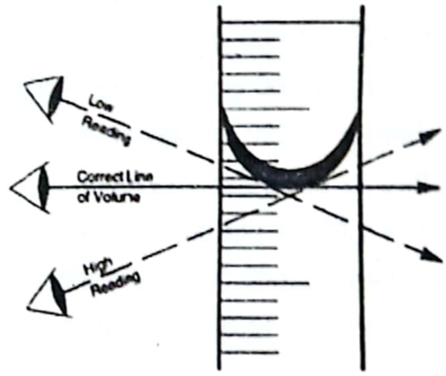
16.4 TOOLS AND TECHNIQUES USED IN EXPERIMENTAL METHODS FOR ACCURACY AND PRECISION OF RESULTS

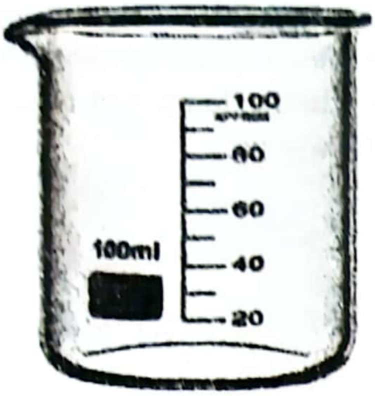
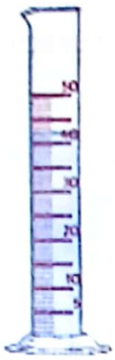
Inaccurate and Imprecise **Inaccurate but Precise** **Accurate and Precise**

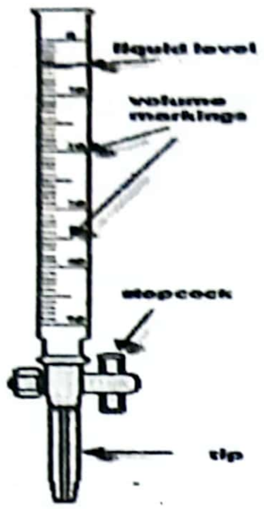
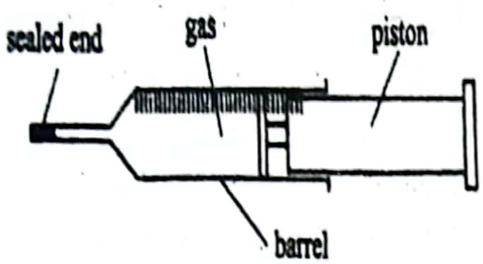
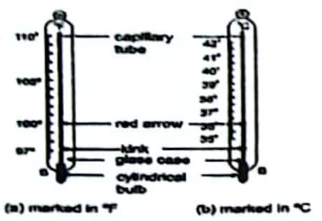
When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true value. Precision refers to how close repeated measurements (using the same device) are to each other. In general, within the laboratory, accuracy is a measure of how well your equipment is adjusted. For example, if your balance is not adjusted correctly, you can make very precise, repeated measurements, but the measurements will not represent the true value. Precision, on the other hand, is usually determined by how careful the scientist is in making measurements. If you are careless and spill part of your sample on the way, your measurements in repeated experiments will not be precise even if your balance is accurate. Different tools and techniques can be used to manage accuracy and precision for inherent errors that arise during measurement

1.6.5 TECHNIQUE AND TOOLS FOR ACCURATE MEASUREMENT IN LABORATORY:

An important technique in a chemistry lab is the ability to accurately measure a liquid in a graduated cylinder.

	 Graduated cylinder	 The right way to take reading
When a solution is in the graduated cylinder the liquid sticks to the sides making a 'U' shape called a meniscus. This makes it difficult to read a measuring cylinder accurately	If you are too high, you will read a smaller volume If you are too low, you will read a larger volume	When you get to 'eye level' you will read an accurate volume of water.

<table>
<tr>
<td>Beaker

</td>
<td>Measuring Beakers are tools used to measure the approximate volume of liquids. Beakers are also used for holding samples, stirring, mixing, and heating liquids. Beakers can also be used as a container for reactions and to estimate the volume of liquid.</td>
</tr>
<tr>
<td>Measuring Cylinder
</td>
<td>Measuring cylinders is more accurate than beakers. They should only be used when an approximate volume measurement is required. They are mostly found in sizes of 10 cm^3, 25 cm^3, 50 cm^3 and 100 cm^3</td>
</tr>
<tr>
<td>Volumetric pipettes
</td>
<td>Most pipettes are used to transfer amounts of up to 1-100 ml of liquids. can only measure a single volume, usually 25 cm^3 . However, they can measure this volume very accurately. They are designed to deliver the measured volume when emptied under gravity, so they hold a slightly greater volume than this. They should only be used with a pipette filler</td>
</tr>
</table>

Burette 	Burettes are designed to deliver any volume up to 50 cm^3. They are more accurate than measuring cylinders but less accurate than pipettes. The volume delivered can be deduced by subtracting the initial measurement from the final measurement. They need to be used with a stand, clamp and boss and they are mainly used in titrations.
Gas syringes 	Gas syringes are used to collect and measure the volume of gas. They have a similar accuracy to measuring cylinders.
Laboratory Thermometer 	The thermometer is used to is used for measuring temperatures other than the human body temperature. It ranges from -10˚C to 110˚C.The laboratory thermometers are used for laboratory purposes such as checking the boiling point and freezing points or temperature of other substances.
Analytical balance 	Analytical balances are precision measuring instruments used in quantitative chemical analysis. It is used to determine the mass of solid objects, liquids, powders, and granular substances.

Beam balance 	It is a valuable tool when an accurate measurement of mass is required. Doesn't require any kind of electricity for its operation.
Electronic balance 	Electronic balance is an instrument used in the accurate measurement of the weight of materials. Electronic balance is a significant instrument for laboratories for precise measurement of chemicals that are used in various experiments. Laboratory electronic balance provides digital results of measurement.
Stopwatch 	Stopwatches and timers are instruments used to measure time intervals, which is defined as the elapsed time between two events. Stopwatches show time reading up to 2 decimal places. So, the precision of most stopwatches is 0.01 seconds.

16.6 ADVANTAGES AND DISADVANTAGES OF EXPERIMENTAL METHODS

Experimental methods have become a valuable part of human life to learn about the world around them. Chemists used different tools and techniques to perform experimental procedures. Marie Curie, Robert Boyle, Linus Pauling, Ernest Rutherford, and Antoine Lavoisier, for instance, did various experiments to uncover key concepts of chemistry. The same goes for modern experts, who utilize this scientific method to see if new drugs are effective, discover treatments for illnesses, and discover new gadgets.

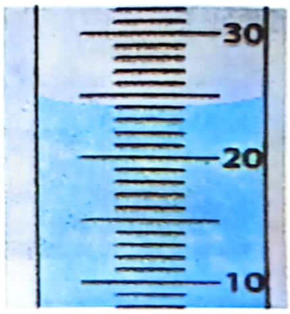
ADVANTAGES	DISADVANTAGES
It gives researchers a high level of control.	It can lead to artificial situations.
It allows researchers to utilize many variations.	It can take a lot of time and money.
It can lead to excellent results.	It can be affected by errors.
It can be used in different fields.	It might not be feasible in some situations.

KEY POINTS

- Physical quantities are measurable quantities.
- Scientific notation is an internationally accepted way of writing numbers in which numbers are recorded using the power of 10 and there is only one non-zero digit before the decimal.
- Scientific measurements use units to quantify and describe the magnitude of something. For example, scientists quantify length in meters.
- Accuracy represents how close a measurement comes to its true value.
- Precision is how close a series of measurements of the same thing are to each other.

REVIEW QUESTIONS

1. Encircle the correct answer.



(i) What is the volume of the liquid in this graduated cylinder?

(a) 23 (b) 24

(c) 25 (d) 22

(ii) How many cubic centimeters (cm^3) are there in 1 decimeter cube (1 dm^3):

(a) 100 cm^3 (b) 1000 cm^3

(c) 10 cm^3 (d) 1 cm^3

(iii) To change SI units by factors of ten into smaller or bigger units they use :

(a) Prefixes (b) symbols

(c) abbreviation (d) ratio

(iv) A distance of 1 kilometer means?

(a) 100 m (b) 1000 cm

(c) 1000 cm (d) 1000 m

(v) One nanometer is equal to:

(a) 10^{-10} m (b) 10^{-9} m

(c) 10^{-8} m (d) 10^{-7} m

(vi) The standard form of 3800000000 is:

(a) 3.8×10^{8} (b) 3.8×10^{9}

(c) 3.8×10^{10} (d) 3.8×10^{11}

(vii) The property of a measuring instrument to give the output very close to the actual value is termed as:

(a) Sensitivity (b) Accuracy

(c) Precision (d) Repeatability

(viii) The standard form of 0.0000000034 is:

(a) 3.4×10^{-8} (b) 3.4×10^{-9}

(c) 3.4×10^{-10} (d) 3.4×10^{-11}

(ix) In SI, the unit of mass is:

(a) kilogram (b) centimetre

(c) kelvin (d) millimetre

(x) In the measuring instruments, the degree of conformity and closeness to the true value is known as:

(a) precision (b) accuracy

(c) sensitivity (d) compatibility

2. Give short answer.

(I) What is system international? Why SI units are standardized for better communication and collaboration.

(ii) In a race, why it is essential to use seconds or minutes as the unit for measurement for recording the time instead of hours?

(iii) Differentiate between accuracy and precision.

(iv) A chemist has a sample of mass 0.003 kilograms. How will he convert this mass to milligrams?

(v) What is the use of prefixes in measurements?

(vi) What are the advantages of using scientific tools like measuring cylinders, stopwatch and thermometers in measurements?

3. How to Calculate the accuracy of measurements?

4. Evaluate how tools for measurements are helpful in performing scientific techniques.

5. How does scientific notation enhance the ability to communicate about extremely large and small numbers? Convince.

6. Why do scientists realize the need for a standardized system of measurement?

PROJECT

Design an experiment to determine density of a liquid.

UNIT 17

SEPARATION TECHNIQUES

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define important terms associated with creating chemical solutions. (some example include: a) solvent as a substance that dissolves a solute. b) solute as a substance that is dissolved in a solvent. c) solution as a mixture of one or more solutes dissolved in a solvent. d) saturated solution as a solution containing the maximum concentration of a solute dissolved in the solvent at a specific temperature. e) residue as a substance that remains after evaporation, distillation, filtration, or any similar process. f) filtrate as a liquid or solution that has passed through a filter.
- Explain methods of separation and purification (some examples includes: a) using a suitable solvent. b) filtration. c) crystallization d) simple distillation e) fractional distillation.
- Suggest suitable separation and purification techniques, given information about the substances involved, and their usage in daily life.
- Identify substances and assess their purity using melting point and boiling point information.

INTRODUCTION

Most of the time the substances that we see around us are not in their pure form. They are a mixture of two or more substances. *Recall* that a mixture contains elements and/or compounds which are not chemically combined together. Interestingly, mixtures tend to also come in different forms. Therefore, there are several types of separation techniques that are used in segregating a mixture of substances.

Consider a problem that people in the 1840's faced. As you know, gold had always been known for its value and therefore was very precious. However, it was hard to come by as it was mixed in the earth and was not easy to separate from all the dirt. *Can you think of any way to separate the gold from the dirt?*

The approach that the people of the time took was to place the soil and gold mixture in a pan and cover with water. The mixture was stirred, the soil would dissolve in the water and the gold would not. After thorough mixing, the pan was gently swirled to remove dissolved material while the heavier gold settled to the bottom of the pan. The gold was then separated from the mixture of soil and water. Through separation, the scientists were able to remove all the unwanted materials, in this case the dirt and water, and obtain the useful components, gold. This is just one method of separating mixtures.

Not only the separation of mixtures is helpful to obtain the useful and valuable components of the mixture, but we are also able to find out the properties of the known/unknown substances from mixtures and possibly use them for the production of useful substances such as medicines.

17.1 METHODS OF SEPARATION

The mixture is composed of two or more substances that can be present in varying amounts and can be physically separated by using methods that use physical properties to separate the components of the mixture. The appropriate method to use for separating the components in a mixture depends on the physical states. Separations exploit differences in chemical properties or their physical properties. For instance, size, shape, mass, density, or chemical affinity between the constituents of a mixture. Some of the common methods of the separation are **filtration, evaporation, distillation, sedimentation and decantation, magnetic separation, centrifugation.**

A solution is a mixture that consists of a solute and a solvent. The solute is a solid that is dissolved in the solvent (which is a liquid) and together they make up a solution. Ask yourself, in a brine solution consisting of salt and water, which is the solute and which is the solvent?

The idea behind the separation techniques of mixtures is to separate the solute and the solvent from each other as they may be much more useful independently. Like in the example above, we separated gold from the dirt and the gold was useful for us.

17.2 EVAPORATION

Evaporation is a separation method use to separate components of a mixture in which a solid is dissolved in a liquid, where the liquid changes into the gaseous phase and solid is left behind. The application is based on the fact that solids do not vaporize easily, whereas liquids do. Can you think of any mixture that can be separated through this technique?

In an enclosed space, a liquid will continue to vaporize until air saturation is achieved. In practically, only a small fraction of the total molecules possess enough heat energy required to vaporize the liquid.

Mixture of a solid dissolved in a liquid

Only a solid remains

17.2.1 Factors affecting rate of evaporation

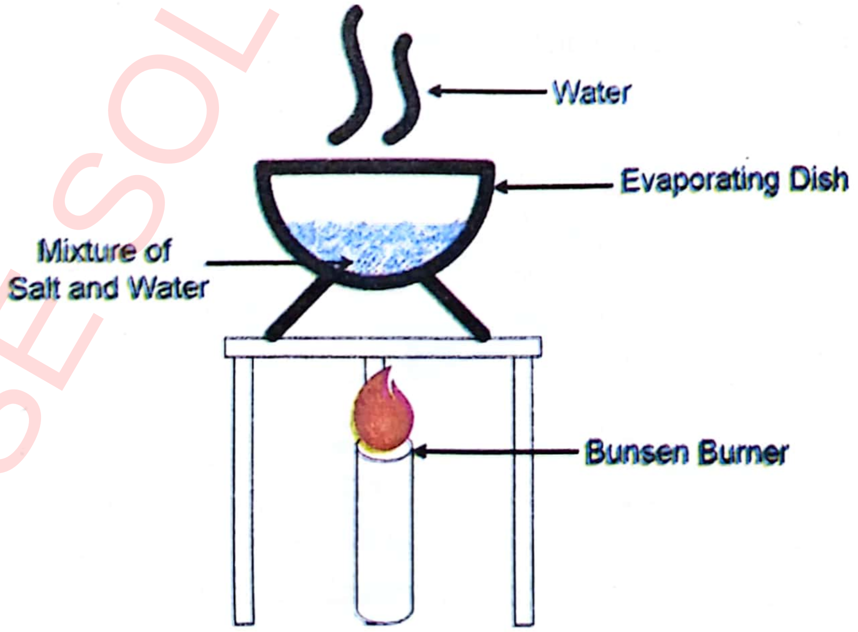
Pressure: An increase in pressure pushes gas particles close together. The force between the particles increases making it difficult for the liquid to convert to gaseous form.

Temperature: On increasing the temperature of the substance, the material heats up and the constituent particles begin moving with greater kinetic energy. This leads to an increase in the rate of evaporation. For example, warm water boils faster than cold water.

Surface area of the substance: A substance with a larger surface area will contain more surface molecules per unit of volume, it triggers the potential of the particles to escape, thereby increasing the rate of evaporation.

Inter-molecular forces: Stronger inter-molecular forces between the molecules of the liquid require greater amounts of energy to break and turn into a gas. Therefore, substances with stronger intermolecular forces have lower rate of evaporation. An indicator of the intermolecular forces is provided by the enthalpy of vaporization.

Flow rate of the atmosphere: Increase in the flow of air which is unsaturated, for instance, fresh air, will lead to the increase in the rate of evaporation.



17.2.2 Applications of Evaporation

From the explanation above, we can deduce that evaporation helps to separate a volatile component from a non-volatile one.

Can you think of any examples?

- seawater can be evaporated to produce drinking water
- Dye from ink: ink is mixture of water and dye
- Dairy producers also use the technique of evaporation to use it to dry lactose into powders.
- Beverage producers can create concentrates, for instance, juice concentrators by reducing water content through evaporation.
- Water containing minerals can be evaporated to extract the metals and minerals which simulates the chemical processes.

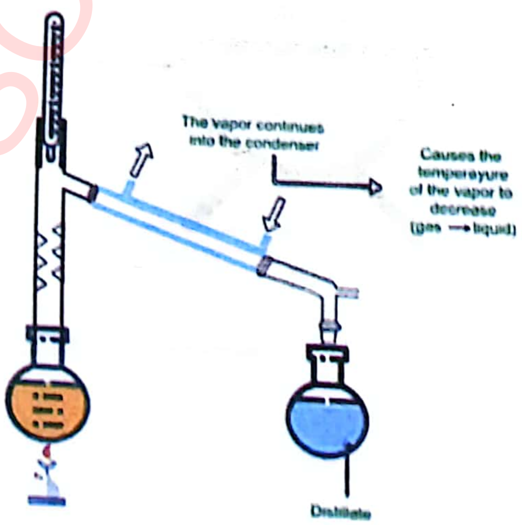
17.3 DISTILLATION

Distillation is a separation technique used to separate components of a liquid mixture by a process of heating and cooling, which exploits the differences in the volatility of each of the components. The liquid that is obtained by performing the condensation of vapor is called the distillate. Distillation is carried out at the solvent's boiling points. Boiling takes place when the vapor pressure is equivalent to the atmospheric pressure.

Separation of the components of the mixture takes place better in case of the higher relative volatility of a liquid. On supplying heat to the liquid, the vapor boils and then condensation takes place.

17.3.1 Distillation procedure:

1. The round bottom flask contains the liquid mixture which is heated to a vigorous boil.
2. The component with the lower boiling point will change into its gaseous state.
3. On coming in contact with the water-cooled condenser, the gas will condense.
4. The condensed gas will trickle down into the graduated cylinder where the chemist can collect the final distilled liquid.
5. The other liquid component remains in the round bottom flask.



17.3.2 Types of Distillation

1. Simple distillation.
2. Fractional distillation.
3. Steam distillation.

Fractional distillation.

This technique works in a very similar way to simple distillation, but fractionating column is used to get a temperature gradient, cooler at the top and hotter at the bottom. This allows careful control of the temperature at which the distillate is being collected, allowing different liquids in a mixture to be separated, each turning into a vapor and being cooled and condensed in the condenser at their own individual boiling points.

Steam distillation.

Steam distillation is a separation process that consists in distilling water together with other volatile and non-volatile components.

Vacuum distillation.

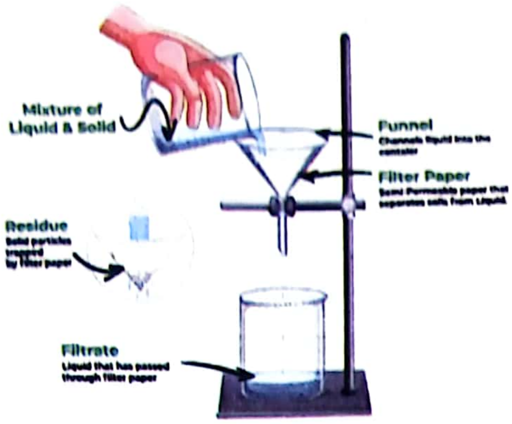
Vacuum distillation or Distillation under reduced pressure is a type of distillation performed under reduced pressure, which allows the purification of compounds which are not readily distilled at ambient pressures or simply to save time or energy. This technique separates compounds based on differences in their boiling points

17.3.3 Application:

- There is a piece of evidence that humans were using the distillation process since 3000 BC (approximately). This process is used to make distilled water.
- Distillation plays an important role in many water purification techniques. Many desalination plants incorporate this method in order to obtain drinking water from seawater.
- This process is also used to refine the crude oil and to purify the alcohol.
- The process of steam distillation is used to obtain essential oils and herbal distillates from several aromatic flowers/herbs.
- Distilled water has numerous applications, such as in lead-acid batteries and low-volume humidifiers.
- Many fermented products such as alcoholic beverages are purified with the help of this method.
- Many perfumes and food flavorings are obtained from herbs and plants via distillation.
- Air can be separated into nitrogen, oxygen, and argon by employing the process of cryogenic distillation.
- Distillation is also employed on an industrial scale to purify the liquid products obtained from chemical synthesis.

17.4 FILTRATION

Filtration is a separation technique used to separate the components of a mixture containing an undissolved solid in a liquid. The exact method used depends on the purpose of the filtration, whether it is for the isolation of a solid from a mixture or removal of impurities from a mixture.



17.4.1 Filtration Process

The principle of filtration is based on the difference in sizes between the particles. Filtration separates the solid matter from a liquid mixture using a filter medium for example filter paper or filtration membrane that allows the liquid to pass through it and block the solid particles on the other side of it. The pores of the filter medium are bigger as compared to the solvent particle but are smaller compared to the solute particle, so they allow only solvent particles to pass through them the solute particle is left on the other side making the solution free from the insoluble solute. The mechanism of "filtration" can be mechanical, biological, or physical. Filtration may be done cold or hot, using gravity or applying vacuum, using a Buchner or Hirsch funnel or a simple glass funnel.

As the process of filtration uses the concept of difference in the particles. We have two types of mixtures

- Heterogenous Mixture
- Homogenous Mixture

We define a heterogeneous mixture as a mixture in which the solute particle is not evenly distributed in the solvent phase. Such as sand in the water is an example of a heterogeneous mixture. On the other hand, a homogeneous mixture is a mixture in which the solute particles are evenly distributed in the solvent. A brine solution is an example of a homogenous mixture. Homogenous mixtures are also called solutions.

Solution

A homogeneous mixture of two or more components is called a solution. Basically, solution has two components

a) Solute: The component of solution which is present in smaller quantity is called solute for example salt, in salt solution.

b) Solvent: The component of solution which is present in larger quantity is called solvent for example water, in salt solution.

Various types of processes are used for filtration. In general, the filtration process uses the filtration membrane that allows only a specific size particle to move through it making the solution at one end free from the insoluble solute particle. Insoluble solid substances left behind are called residues. We can understand this by using the example, suppose we have to filter sand

and water then we use a muslin cloth at the mouth of the container containing the sand in water. We allow the solution to pass through the muslin cloth which only allows water may pass through it separating the sand from the sand and water solution. We can also use filtration paper in place of muslin cloth as the filtration membrane.

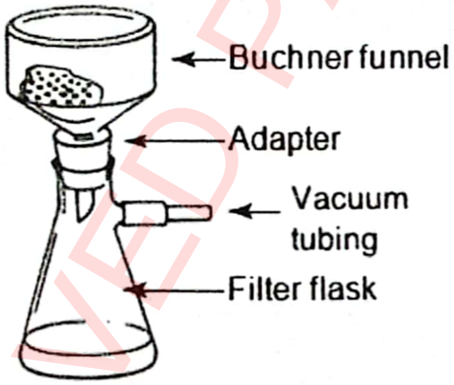
7.4.2 Filtration by filter paper

Filtration by filter paper is also called gravity filtration. Filter paper is a semi-permeable barrier placed perpendicular to a liquid or air flow. Filter paper can have pore sizes ranging from small to large to permit slow to fast filtering

17.4.3 Filtration by vacuum (or Suction) Filtration

Vacuum filtration uses a Buchner funnel. A Buchner funnel is a flat bottomed, porous, circular porcelain bowl with a short stem. The stem is fitted with a rubber stopper and inserted in the mouth of a side arm filter flask. Circular filter paper, the same diameter as the bowl, is placed on the flat bottom and wetted with the appropriate solvent to create a seal before starting the filtration.

Can you think of some examples of filtration in your daily life? Some are listed below.



17.4.4 Filtration in daily life:

- When making coffee, hot water is filtered through the ground coffee and a coffee bean. The filtrate is the coffee in liquid form. Using a tea ball or a tea bag (paper filter) for tea is also an example of filtration.
- A kidney is a biological filter that is used to filter the human blood. Blood filtering in kidneys is another application of the filtration technique.
- HEPA filters are used by various vacuum cleaners and air conditioners to filter out dust and pollen from the air.
- Filters with particulate-capturing fibers are commonly used in aquariums.
- During mining, belt filters extract precious metals. Filtration techniques are used in the metallurgical process to remove the slag.
- Sand and permeable rock in the ground filtered the water which is then stored in the aquifer and

then used as groundwater.

- The wastewater treatment plant uses the filtration technique to filter the sewage.
- Air filters are used in automobiles and in factories to remove harmful particles from the smoke.
- The treatment of water uses filtration techniques.

17.5 CRYSTALLIZATION

Crystallization is a technique for purification of the substances. It is a technique that separates a solid from its solution. **The process in which the solid dissolves in the liquid arranges itself into well-defined 3-D structure is called crystallization. When any substance undergoes crystallization its molecule arranges itself in a fixed structure at an angle to form a 3-D structure called the crystal and then the crystal is removed from the solution. The crystal is a pure substance and thus the process of crystallization is widely used in the purification of material from its impure solution.**

What is the principle of crystallization?

The principle of crystallization is based on the limited solubility of a compound in a solvent at a certain temperature and pressure. A change of these conditions to a state where the solubility is lower will lead to the formation of a crystalline solid.

Crystallization is a purification technique that involves the following steps:

- When a small amount of solute at a given temperature is added in a solvent, solute dissolves very easily in the solvent.
- If the addition of solute is kept on, a stage is reached when solvent cannot dissolve any more solute.
- At this stage, further added solute remains undissolved and it settles down at the bottom of the container. On the particle level, a saturated solution is the one, in which undissolved solute is in equilibrium with dissolved solute.
- At this stage, dynamic equilibrium is established. Although dissolution and crystallization continue at any temperature, the net amount of dissolved solute remains constant.
- The solution is heated until enough of the solvent (water) has evaporated to make the solution saturated.

- We can tell when the solution is saturated by dipping a glass rod into the solution to remove a drop and seeing if the drop goes cloudy and crystals start to form as it cools.
- Once the solution is saturated, the Bunsen is turned off and the solution is allowed to cool - solubility decreases with temperature, and the solute that can't remain dissolved forms crystals.
- The crystals are removed from the remaining solution by filtering.
- The crystals are dried (in a warm oven, or just left to dry in air)

Science Titbits:

The size of the crystals formed during this process depends on the speed of cooling.

Rapid cooling of the solution will cause a large number of tiny crystals to form.

Slow cooling rates lead to the formation of large crystals.

17.5.1 Application of Crystallization

Crystallization is primarily employed as a separation technique in order to obtain pure crystals of a substance from an impure mixture.

Another important application of crystallization is its use to obtain pure salt from seawater.

Crystallization can also be used to obtain pure alum crystals from an impure alum. In such scenarios, crystallization is known to be more effective than evaporation since it also removes the soluble impurities

17.5.2 Application of separation and purification techniques

Filtration, Crystallization, distillation are suitable separation and purification techniques Filtration can be a physical process or a biological process. The physical application of filtration can be seen routinely all around our daily life–

- Human kidneys are responsible for filtering both the blood and other waste materials that may enter the body, whether through food, drink or medicine. The waste leaves the body as urine.
- coffee filters, automotive filters are commonly used in daily routine.
- Tea staining or coffee filters are examples of filtration Common example of crystallization for separation and purification the salt we get from seawater can have many impurities in it. Hence, the process of crystallization is in use to remove these impurities.
- The crystallization of water to form ice cubes and snow is commonly observed in daily life.
- The crystallization of honey when it is placed in a jar and exposed to suitable conditions.
- The formation of stalagmites and stalactites (especially in caves).
- Have you ever observed as the sugar syrup cools, sugar crystals can begin to form. This is crystallization.

17.5.3 Identification and purity of substances

Melting Point and Boiling point are characteristic parameters of substances which help to identify them. Pure substances have a fixed boiling and melting point. For example, pure water boils at a 100 °C and melts at 0 °C. Impure substances have no fixed melting and boiling points. They can melt and boil at a range of temperatures.

To determine if a substance is pure in school laboratories, we can check the substance's melting or boiling points. Impure substances tend to have a slightly lower melting point and higher boiling point than the pure substance, and a broader range.

Have you ever noticed that in chocolate - milk chocolate melts really easily in your mouth whereas dark chocolate which is usually 70% takes much longer to melt. When water contains dissolved impurity (for example, table salt), it boils at a slightly higher temperature (around 100.5°C).

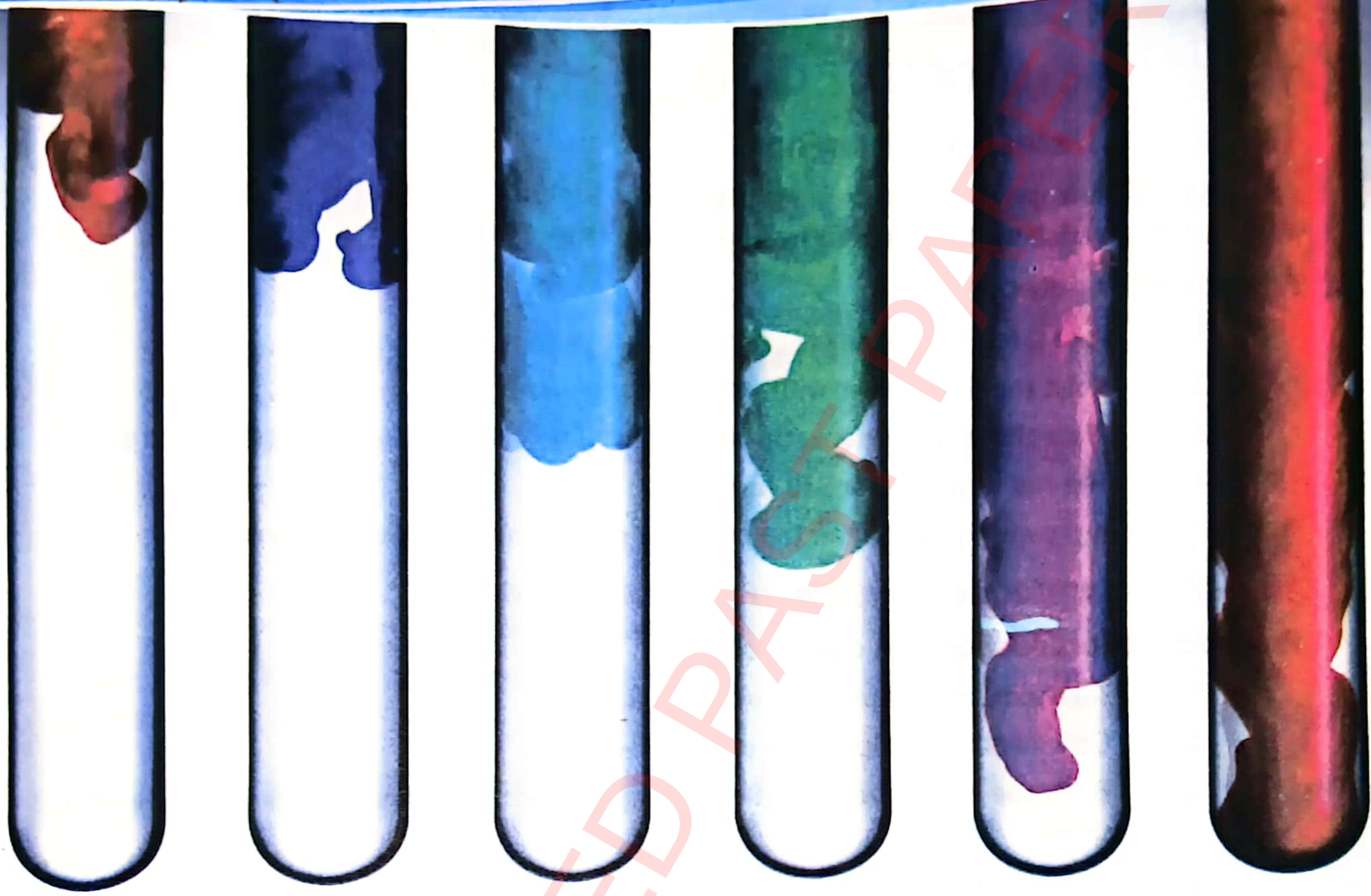
KEY POINTS

- The change of a liquid into its vapours is called evaporation.
- Distillation is a separation technique used to separate components of a liquid mixture by a process of heating and cooling.
- Distillation under reduced pressure is called vacuum distillation.
- Filtration separates the solid matter from a liquid mixture using a filter medium.
- The principle of crystallization is based on the limited solubility of a compound in a solvent at a certain temperature and pressure.

REVIEW QUESTIONS

1. Give short answer.
   (i) Why is it important to separate materials from a mixture?
   (ii) What is distillation
   (iii) What is filtration?
   (iv) What is evaporation?
2. What technique would you use to separate sand from water? There are two possibilities. Mention their names.
3. What technique would you use to separate alcohol from water?
4. What are mixtures?
5. How separation of mixtures of two or more liquids is done. Name the separation technique.
6. How to separate a mixture of two solids?
7. What method will you use to separate a sand and sugar mixture?

UNIT 18



QUALITATIVE ANALYSIS

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Describe tests to identify important gases (Some examples include: a)ammonia, NH_3, using damp red litmus paper b) carbon dioxide, CO_2, using lime water c) chlorine, Cl_2, damp litmus paper d) hydrogen, H_2, using a lighted splint) oxygen, O_2, using a glowing splint e) sulphur dioxide, SO_2, using acidified aqueous potassium manganate(VII)
- Explain the use of a flame test to identify important cations: (Some examples include: a) lithium, Li^+ b) sodium, Na^+, c) potassium, K^+, d) calcium, Ca^{2+}, e) copper, Cu^{2+}, f) barium, Ba^{2+}

INTRODUCTION

Qualitative analysis is those in which one only tells about the nature of a substance and not its quantity. These analysis methods have different reliability. Some of them are very reliable and reach up to confirmatory tests.

18.1 DETECTION OF GASES

18.1.1 Detection of ammonia

Ammonia is a basic gas whose pH is more than 7. It is used for making urea fertilizer. It can be manufactured by Haber's process using hydrogen and nitrogen gases.

Procedure for detection

1. Take a red litmus paper.
2. Moist the filter paper with water spray
3. Place the litmus paper at the mouth of test tube or flask containing ammonia gas.
4. The red litmus paper will turn blue.

Turning the litmus paper blue confirm the presence of any basic gas like ammonia.

18.1.2 Detection of carbon dioxide (CO_2)

Procedure of production and detection

1. Take some marble pieces in a conical flask.
2. Cork it with the help of the cork with two holes.
3. From one hole insert a thistle funnel, and from the other whole a glass tube of u shape.
4. Insert the other end of tube in a test tube containing lime water ($Ca(OH)_2$) aq.
5. Hydrochloric acid in the thistle funnel.

Figure 18.1: Detection of carbon dioxide.

6. Carbon dioxide gas formed will travel through the tube into lime water.
7. $CaCO_3+HCl \rightarrow CaCl_2+CO_2+H_2O$
8. The lime water will turn milky due to the production of calcium carbonate which is insoluble.
9. $CO_2+Ca(OH)_2 \rightarrow CaCO_3+H_2O$

This method is a confirmitory test for Carbon dioxide.

18.1.3 Detection of chlorine gas

Detection procedure

1. Prepare Damp Litmus Paper:

   Take a piece of blue litmus paper and dampen it with distilled water. Blue litmus paper is typically used for acidic gases like chlorine.
2. Expose to the Gas:

   Place the damp blue litmus paper in the test tube of chlorine gas.
3. Observe Colour Change:

   If chlorine gas is present, the blue litmus paper will turn red. This colour change occurs because chlorine gas is acidic and reacts with the red litmus paper, causing it to change from blue to red and than bleaches it to white.

Interestingly blue litmus after turning red turns white as it bleaches the paper, chlorine gas is a bleaching agent.

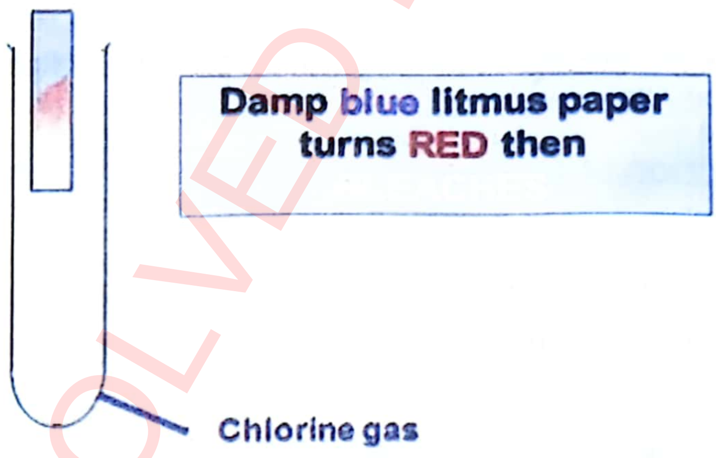


Figure 18.2: Detection of chlorine.

18.1.1 Detection of hydrogen gas using lighted splint (pop reaction)

hydrogen gas can be prepared by the reaction of zinc metal with sulphuric acid. It can be stored in an open inverted test tube. Because of its light weight it does not go out in a vertically inverted test tube.

Procedure:

1. Put some zinc metal pieces in a wolf bottle.
2. Wolf bottle has two openings, from one opening using a cork pass a thistle

funnel, from the other opening pass a U-shaped tubing.

3. Make the apparatus air tight by using wax.
4. Take a test tube to collect hydrogen gas coming from the tubing. And place it vertically inverted over the tubing
5. Pour sulphuric acid from the thistle funnel.
6. $Zn_{(s)}+H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)}+H_{2(g)}$

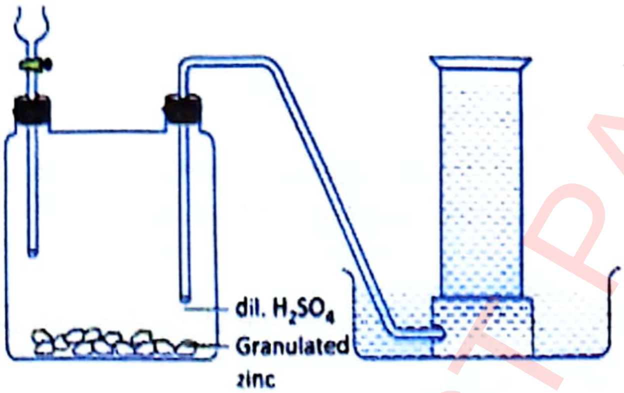


Figure 18.3: Formation of hydrogen.

7. The test tube will be filled with the hydrogen gas.
8. Bring aap burning splint after the test tube is filled with hydrogen,
9. A popping voice with a flame is observed which will confirm the presence of hydrogen gas. $2H_{2(g)}+O_{2(g)} \rightarrow 2H_2O_{(l)}$

This happens as a reaction of hydrogen with oxygen in the presence of flame and the formation of water vapours.

18.1.2 Detection of oxygen by using a glowing splint

Materials required

1. A glowing wooden splint (a small stick or piece of wood)
2. A source of oxygen (e.g., oxygen gas tank or hydrogen peroxide solution)
3. A test tube
4. A match stick

Procedure:

Ensure you are in a well-ventilated area and follow safety precautions when handling flammable materials.

Prepare your source of oxygen:

use hydrogen peroxide, pour a small amount into the test tube, it generates oxygen gas automatically.

Ignite the wooden splint using a matchstick or lighter until it's glowing at the tip. Carefully insert the glowing splint into the test tube containing the oxygen gas.

Observe what happens to the splint inside the test tube, if oxygen is present, the splint will burst into flames, burning brightly.

Oxygen gas helps in combustion, thus it will suddenly enflame the vanishing splint.

18.1.6 Detection of Sulphur dioxide

SO2 is a very good reducing agent, it can be detected by using an oxidizing agent like Potassium permanganate.

Procedure

1. Prepare a solution of potassium permanganate ($KMnO_4$) in water it will be of bright purple colour.
2. Pass the gas sample containing SO_2 through the $KMnO_4$ solution.
3. The following reaction occurs in which purple colour of $KMnO_4$ will be discharged. It shall indicate the presence of SO_2 gas.
4. $2KMnO_4 + 3SO_2 + H_2O \rightarrow 2MnO_2 + 2K_2SO_4 + 2H_2SO_4$

This method is often used in analytical chemistry and environmental monitoring to detect and measure SO_2 levels in gases.

18.2 DETECTION OF METAL CATIONS BY FLAME TEST

A Bunsen burner flame has two parts the upper part is called oxidizing flame and the lower part is called reducing flame.

Bunsen flame has enough energy to excite the electrons of alkali and alkaline earth metal atoms.

Alkali metals have low ionization potentials that's why there atoms get their electrons excited by the energy of visible wavelength of light. The same energy is released when the electron comes to the ground state showing the colour of flame, as this flame contain the energy of visible wavelength of light it possesses a characteristic flame colour.

How is flame test performed?

1. Sample Preparation: A small amount of the sample e.g. salt, oxide or any other compound containing the metal ions is usually dissolved in water to create a solution or paste.
2. Clean Wire Loop: A clean, non-reactive wire loop of platinum or even a clean glass rod is used to hold a small amount of the sample solution.
3. Heating: The wire loop with the sample is then introduced into the flame of a Bunsen burner. The heat causes the metal ions in the sample to become excited.
4. Observation: As the metal ions return to their ground state from the excited state, they emit light in the form of characteristic coloured flames. The colour of the flame is then observed and compared to a reference chart to identify the metal ion present.

   Sodium gives golden yellow, lithium and strontium give crimson red, Potassium gives purple, Cesium gives blue, barium gives green and copper gives bluish green flame colours.

Figure 18.4: Flame of different metallic cations.

KEY POINTS

- Qualitative analysis is those in which one only tells about the nature of a substance and not its quantity
- Ammonia gas turns red litmus paper blue
- Carbon dioxide turns lime water milky
- Chlorine changes blue litmus paper red and then bleaches to colourless
- Hydrogen burns with pop sound
- Oxygen helps in combustion
- Sulphur dioxide discharges the purple colour of $KMnO_4$
- Sodium gives golden yellow flame, lithium and strontium give crimson red, Potassium gives purple, Cesium gives blue, barium gives green and copper gives bluish green flame

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Ammonia is a gas
(a) Acidic (b) Basic
(c) Neutral (d) Amphoteric

(ii) Which gas helps in combustion process?
(a) Oxygen (b) Nitrogen
(c) Sulphur dioxide (d) Carbon dioxide

(iii) Which guess turns the lime water, milky?
(a) Carbon monoxide (b) Carbon dioxide
(c) Sulphur dioxide (d) Oxygen

(iv) Electrons of alkali metals excite by absorbing the light in:
(a) Visible wavelength (b) UV wavelength
(c) IR wavelength (d) Radio wavelength

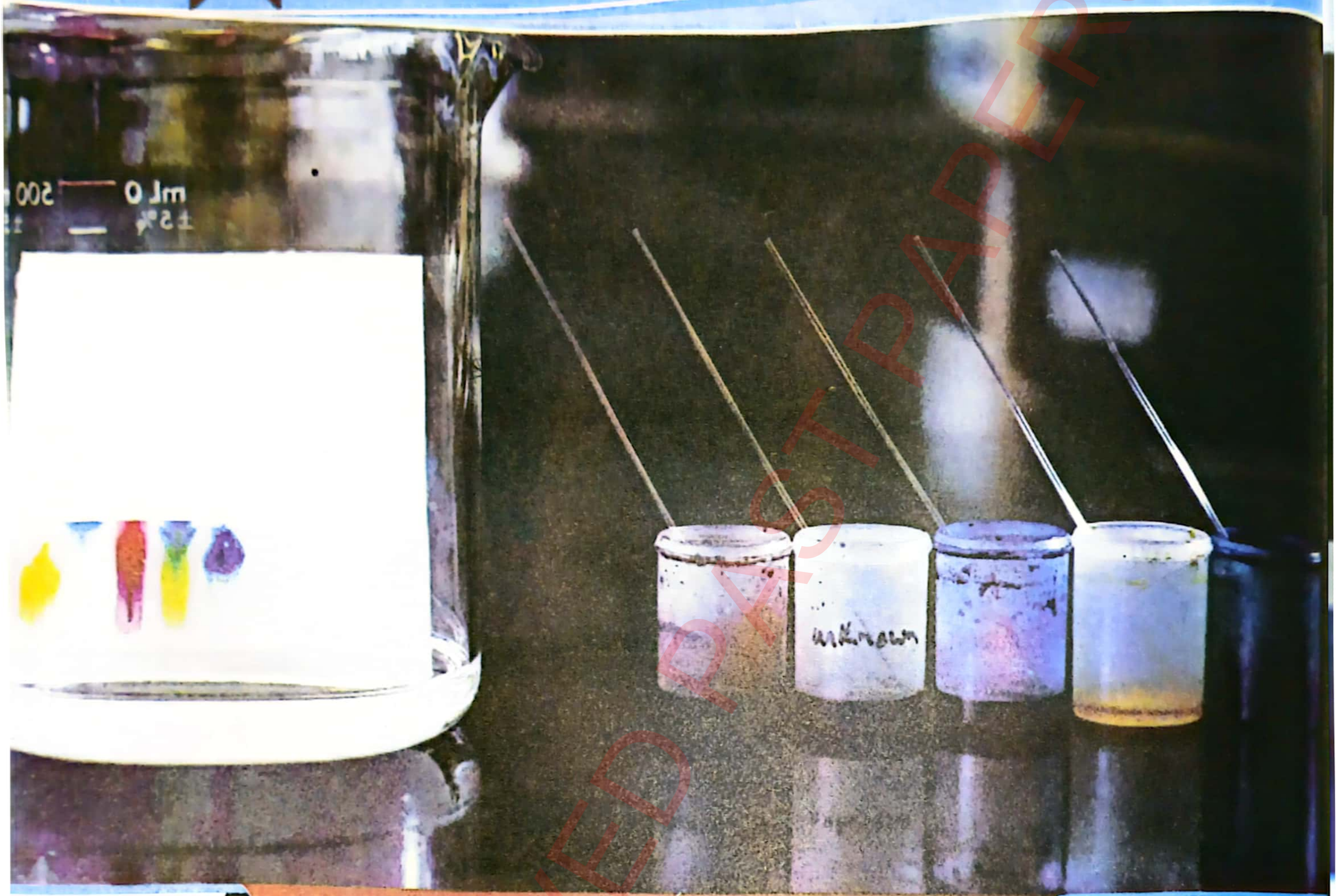
(v) The colour imparted by the flame of sodium metal is:
(a) Blue (b) Green
(c) Golden yellow (d) Purple

2. Give short answer.

(i) How is ammonia detected by a litmus paper?
(ii) How can you identify carbon dioxide?
(iii) How can you detect hydrogen gas?
(iv) How can you detect sulphur dioxide?
(v) Differentiate between oxidizing and reducing flame?

3. What is the difference between qualitative and quantitative analysis?
4. What is the origin of flame colour?
5. Why does hydrogen gas produce a popping voice when it is exposed to the flame?
6. Why the litmus paper turns white at the end after converting red.

UNIT 19



CHROMATOGRAPHY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Describe how paper chromatography is used to separate mixture of soluble substances, using a suitable solvent.
- Describe the use of locating agents when separating mixtures containing colourless substances. (F or context, knowledge of specific locating agents is not required)
- Interpret simple chromatograms (for context, students should identify:
  - a) unknown substance with known substances
  - b) pure and impure substances.
- State and use the equation for R_f.

INTRODUCTION

The word chromatography has been originated from two words one is Chroma means colour the other graphein means writing.

Chromatography has been originated as a very strong analytical tool in the analytical chemistry. This is not only used to analyse a compounds, it is used for separation of the compounds as well. Nowadays this technique can be coupled with other techniques like mass spectrometry.

19.1 PRINCIPLE OF CHROMATOGRAPHY

Chromatography works on the principle of partitioning where the components of a mixture distribute themselves between two phases one is called stationary phase and the other one is said to be mobile phase.

Stationary phase:

Stationary phase in chromatography is a solid phase or a liquid phase coated on the surface of a solid phase. Like in paper chromatography water molecules are adsorbed on paper material cellulose.

Mobile phase:

It is a liquid or gas that moves through or over the stationary phase.

As the mixture is carried by the mobile phase through the stationary phase, each component interacts differently with the two phases, leading to different migration rates. This results in the separation of the components based on their affinity for the stationary phase relative to the mobile phase.

19.2 Paper chromatography

In paper chromatography the stationery phase is the water absorbed on filter paper. The mobile phase can be a solvent or a mixture of solvents. A mixture of inks or some metal ions can be separated with the help of paper chromatography.

Procedure

1. Take 1 dm^3 measuring cylinder fitted with the rubber bung carrying a glass hook at the lower end
2. Pour 25 cm^3 of water and ethanol solvent mixture (6:4) in it and close it. Allow it to stand for 15 minutes.
3. Take a strip of paper (what Mann filter paper number one) 2.5 cm wide and lengthier then the length of the cylinder.
4. Mark a line 5 centimetre from one end its mid point with pencil.
5. Prepare a mixture of various inks like blue, black, green, red etc by taking two drops of each in a clean China dish. Mix them thoroughly.
6. Apply a very tiny drop of this mixture at the midpoint by means of capillary tube.
7. Dry the strip in the air for about 15 minutes and suspend it by the glass hook with the impregnated end dipping into the solvent to a depth of 5-6 mm.

8. When the solvent front has risen for about an hour or the solvent front is 2-3cm below the upper end of the filter paper.
9. Remove the paper strip and dry it in air. It is seen that every ink gives coloured bands at different regular intervals.
10. Measure the distance of these coloured bands from the baseline and the distance travelled by the solvent front from the baseline.
11. alculate the Rf values for various inks by the formula given. In this way we can separate a mixture of various inks by chromatography.

$$R_f = \frac{\text{Distance traveled by solute}}{\text{Distance traveled by solvent}}$$

R_f factor(retardation factor) is a measure of affinity of a component with mobile phase or stationary phase. If R_f value is greater it means that the component is more affine towards the mobile phase than to stationary phase.

Locating agent

A locating agent is a substance or a solution which is used to locate the components which have been separated by chromatography but are not visible due to being colourless. For example, a mixture of amino acids can be separated by paper chromatography or by thin layer chromatography (TLC) but they are not visible, to locate them ninhydrin is sprayed on the chromatogram. Ninhydrin reacts with amino Acids and develops into a colourful chromatogram.

Reading a chromatogram

Measure the distance each compound spot travelled from the starting point (origin) to its current position. This is known as the R_f (Retention Factor) value and is calculated as the distance travelled by the compound divided by the total distance travelled by the solvent. Compare the R_f values of the separated compounds to known standards data to identify the components within your chromatogram.

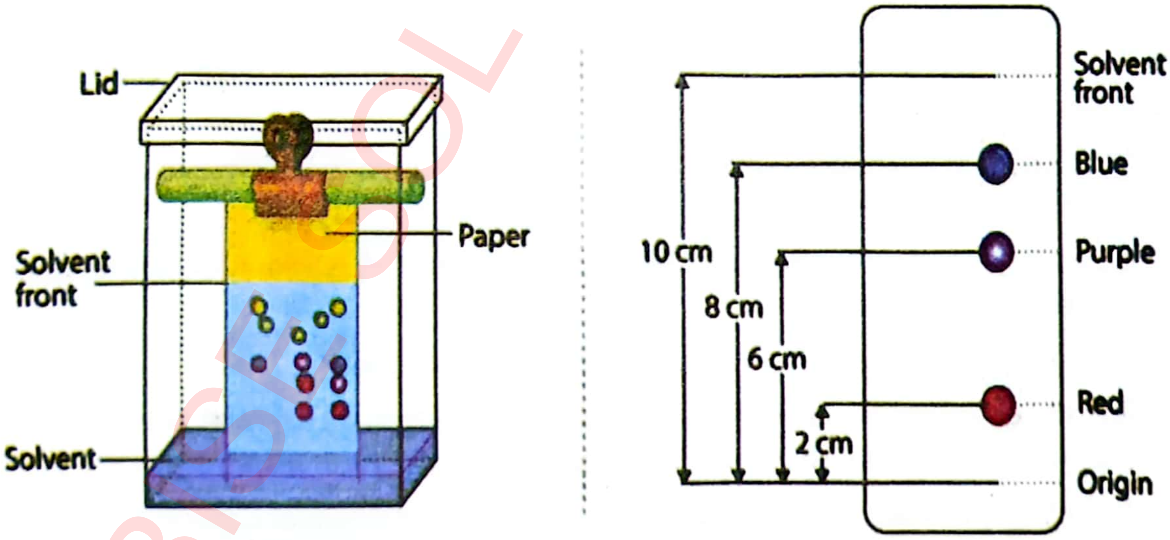


Figure 19.1: Paper chromatography

For example, in the above given chromatogram solvent front travelled 10cm and the blue coloured component travelled 8 cm, so the R_f is:

R_f (blue) = 8/10 = 0.8

R_f (purple) = 6/10 = 0.6

R_f (red). = 2/10 = 0.2

Blue component has the greatest affinity with the mobile phase and the red has the greatest affinity with stationary phase.

Identification of an unknown substance

A paper chromatogram is used to identify an unknown substance by comparing it with known substances. Two substances are likely to be same, if their R_f Value is same. Therefore, R_f value of an unknown substance is determined. This R_f value is then cross-matched with already known R_f values to identify the potential substance.

A pure substance produces only one spot on the paper chromatogram. An impure substance can produce more than one spots depending on the compounds present in it.

KEY POINTS

- Chromatography works on the principle of partitioning where the components of a mixture distribute themselves between two phases one is called stationary phase and the other one is said to be mobile phase.
- In paper chromatography the stationery phase is the water absorb on filter paper
- A locating agent is a substance or a solution which is used to locate the components which have been separated by chromatography but are not visible due to being colourless
- Rf (Retention Factor) value is calculated as the distance travelled by the compound divided by the total distance travelled by the solvent

REVIEW QUESTIONS

1. Give short answer.
   (i) Define chromatography
   (ii) Define locating agent
   (iii) What is R_f value?
   (iv) Define paper chromatography
   (v) What do you mean by paper chromatogram?
2. Explain in detail paper chromatography.
3. How can you identify an unknown substance by chromatography?
4. Differentiate between stationary and mobile phase.

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GLOSSARY

Allotropes are different physical forms of a pure element.

Alkanes the simplest hydrocarbons in which all carbon-carbon bonds are single.

Alkenes The family of hydrocarbons which contain one carbon-carbon duble bond.

Alkynes unsaturated hydrocarbons with a carbon-carbon triple bond.

Arrhenius Acid a substance that produces H^+ ions in water.

Arrhenious base a substance that produces OH^- ions in water.

Base a substance that can accept a proton from an acid.

Bronsted lowry acid a substance that can donate a proton.

Bronsted lowry base a substance that can accept a proton

Alpha particle is doubly positively charged helium nuclei.

amu is the abbreviation for atomic mass unit. It is exactly $1/12^{th}$ mass of one atom of carbon-12.

Anion is an atom or group of atoms having a negative charge

Atmosphere is the unit of pressure of air at the earth's surface.

Atom is the smallest particle of an element to retain the properties of the element.

Atomic mass is the exact mass of one atom of an element.

Atomic number is the number of protons in the nucleus of an atom.

Bonding is the term used to describe ways in which atoms are held together in elements and compound.

Cehmical equilibrium -a state of a chemical reaction at which rate of forward reaction equals the rate of reverse reaction.

Cation is an atom or group of atoms having a positive charge.

Chemical bond is the force that holds two or more atoms together in a substance.

Chemistry is the study of matter, its composition, properties and the interaction of atoms as they react to form new substances.

Colloid is a heterogeneous mixture of a substance dispersed through a medium.

Compound is a substance made up of elements in fixed ratio.

Concentration is the measure of the amount of solute dissolved in a solution.

Condensation is the change of a gas into a liquid.

Conductor is a substance that can conduct electric current.

Covalent bond is a bond formed by the mutual sharing of electrons between two atoms.

Distillation is the process for purifying a liquid by evaporating it, condensing the vapours and collecting the liquid in a clean container.

Double bond is a covalent bond formed by two pairs of shared electrons.

Electrolyte is a substance which in fused form or in solution form allows current to flow through it.

Electron is the sub-atomic negatively charged particle in an atom.

Electron configuration is the arrangement of electrons in orbitals.

Electron shells are sets of energy levels corresponding to a single value of n.

Electronegativity is the power of an atom to attract electrons to itself.

Element is the pure substance in which all the atoms have same atomic number.

Energy levels	are the discrete energies of electron in atom
Enhanced greenhouse effect	an upset in the natural balance of the concentration of greenhouse gases in the atmosphere.
Evaporation	is the escape of molecules of a liquid through the surface.
Formula mass	is the sum of atomic masses of all the atoms present in the compound. It can also be defined as the mass of one formula unit or molecule in amu or mass of one mole of the compound in grams.
Formula unit	is the simplest formula of an ionic compound.
Freezing point	is the temperature at which a liquid changes to a solid.
Functional groups	are groups of atoms in an organic molecule which give particular properties.
Gas	is one of the physical states of matter that has neither any definite shape nor fixed volume.
Gram formula mass	is the mass of 1mole of a compound or element in grams.
Group	refers to the vertical column on the periodic table.
Halogens	are the elements present in VIIA group on the periodic table.
Hydrochloric acid	is an aqueous solution of hydrogen chloride.
Hydrogen bond	is an intermolecular force due to the interaction of lone pair present on N, O or F atom and the partially positively charged hydrogen of another molecule attached with N, O or F.
Inorganic	refers to a compound made up primarily of atoms other than carbon.
Intermolecular forces	are the attractive forces between the molecules.
Ion	is an atom or a molecule possessing an electric charge.
Ionic bond	is the bond formed by the attraction of oppositely charged ions.
Isotopes	are atoms of an element that differ in the number of neutrons.
Liquid	is one of the states of matter that has definite volume but no definite shape.
Lone pair	is a pair of non-bonding valence electron or a pair of electron that are not shared in covalent bonding.
Main group elements	are elements belonging to sub-groupA on the periodic table.
Mass number	is the sum of the number of protons and neutrons in an atom.
Matter	is anything that occupies space and has mass
Melting point	is the temperature at which a substance melts.
Metals	are elements that are good conductor of electricity and heat and tends to form positive ions.
Molar mass	is the mass of one mole of a compound or element in grams.
Molarity	is the number of moles of solute per dm^3 of solution.
Mole	is Avogadro's number of atoms, molecules or formula units. It can also be defined as the atomic mass, formula mass or molecular mass of a substance expressed in grams.
Molecular mass	is the sum of atomic masses of all the atoms present in a molecule.
Molecule	is a group of atoms held by a chemical bond and behaves as an independent unit.
Neutron	is a sub-atomic neutral particle in the nucleus of an atom.
Non-metal	is an element that is a bad conductor of electricity and tends to form negative ion.
Nucleus	is the positively charged center of an atom.
Octet	is eight electron configurations in the valence shell.
Organic chemistry	is the study of carbon compound or the study of hydrocarbons and

	their derivatives.
Organic compounds	are compounds mainly composed of carbon atoms except bicarbonates, carbonates, oxides and carbides.
Oxidation	is the loss of electron by an element.
Oxidation state	is the actual charge on an ion or apparent charge on an atom in a molecule.
Oxidizing agent	is the reactant containing the element that is reduced.
Period	refers to the horizontal row in the periodic table.
Periodic table	is s systematic arrangement of elements based on electron configuration and periodic behaviour.
Polar covalent bond	is a bond in which electrons are shared unequally between two atoms. Thus one atom bears partial positive charge and the other a partial negative charge.
Proton	is the positively charged sub-atomic particle in the nucleus of an atom.
Reactant	is a substance consumed in a chemical reaction.
Redox reaction	is an oxidation and reduction reaction.
Reduction	is the gain of electron by an atom.
Reducing agent	is the reactant containing the element that is oxidized.
Refining	a process of separation of crude oil into useful fractions.
Solid	is one of the three states of matter, it has a definite shape and definite volume.
Solubility	is the amount of solute that dissolved in 100g of water at a particular temperature.
Solute	is the component of solution that is present in relatively small amount.
Solvent	is the component of solution that is present in relatively larger amount.
S-orbial	is a spherical shape orbital.
Structural formula	shows the arrangements of atoms in a molecule
Sublimation	is the conversion of a solid directly into a gas.
Technology	is the application of scientific knowledge to serve some practical purpose.
Triple bond	is a covalent bond involving the sharing of three pairs of electrons.
Unsaturated hydrocarbons	molecules containing only carbon and hydrogen in which at least two carbon atoms form double or a triple bond.
Universal indicator	an indicator that changes colour at each pH value.
Valence electrons	are the electron in the outer most energy level of the atom.
Valence shell	is the outermost energy level of the atom.
Vapour pressure	is the pressure exerted by the vapours of a liquid in equilibrium with its liquid at a particular temperature.
Zero group	refers to VIIIA group on the periodic table

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تو نشانِ عزمِ عالی شان ارضِ پاکستان
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شاد باد منزلِ مُراد!

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