Model Textbook of

CHEMISTRY

Grade 9

Based on National Curriculum of Pakistan 2022-23

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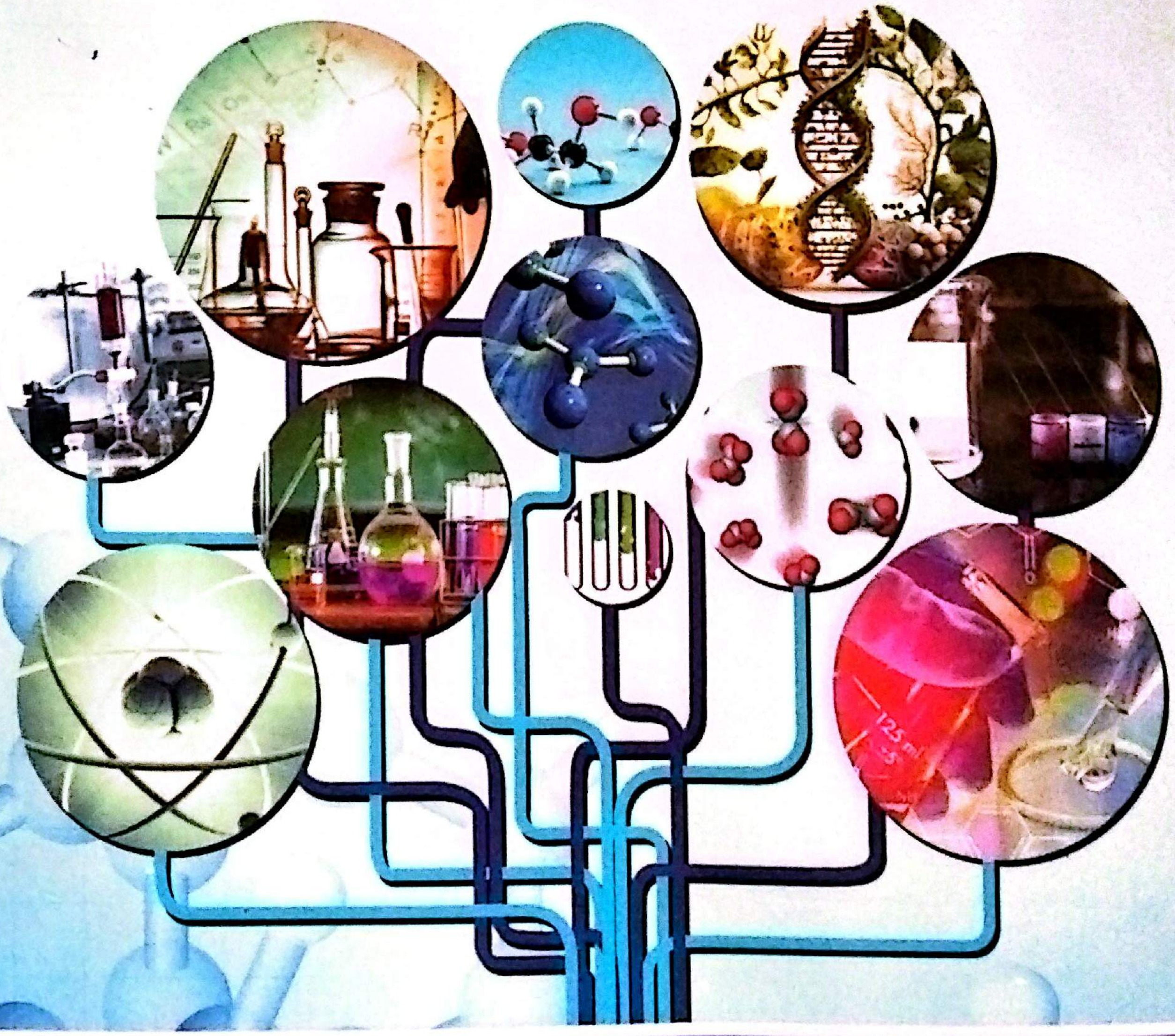

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Based on National Curriculum of Pakistan 2022-23

Arfah Nadeem
9th-A

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National Curriculum Council
Ministry of Federal Education and Professional Training

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Approved by the National Curriculum Council (NCC), Ministry of Federal Education and Professional Training, Islamabad
vide letter No. F.No 1-1(2024)-NCC/DEA/Dir/English, Dated: March 08, 2024

A Model Textbook of Chemistry for Grade 9

Based on National Curriculum of Pakistan (NCP) 2022-23

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Printed in Pakistan

First Edition: First Impression: March 2024 | Pages: 258 | Quantity: 157000

First Edition : Second Impression: May 2024 | Pages: 258 | Quantity: 65000

Second Edition : First Impression: January 2025 | Pages: 254 | Quantity: 89050

Second Edition : Second Impression: January 2025 | Pages: 254 | Quantity: 17031

Second Edition : Third Impression: April 2025 | Pages: 254 | Quantity: 10000

Second Edition : Forth Impression: February 2026 | Pages: 254 | Quantity: 57000

Price: PKR 625/- , **Code:** 978-969-37-1580-0, **ISBN:** 571-479

Printer: GPS Printing Press Lahore

For details on additional publications from the National Book Foundation, please visit our website at www.nbf.org.pk

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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. Kamran Jahangir

Managing Director

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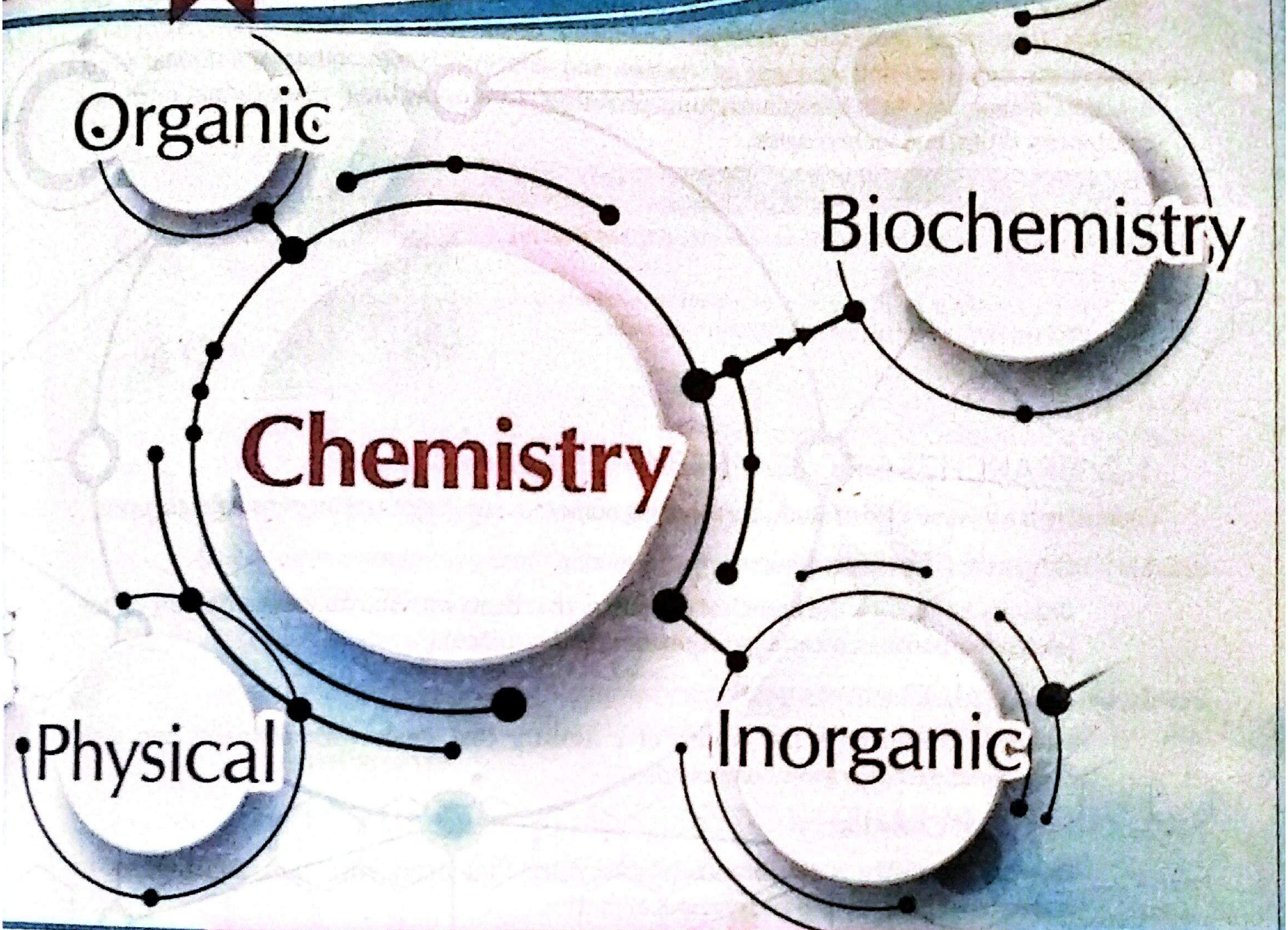
UNIT 01



NATURE OF CHEMISTRY IN SCIENCE

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

2. **Inorganic Chemistry**

   Inorganic chemistry is the 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 the 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 that 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 the 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.

SCIENCE
Science is the systematic study to explore the natural world. Science intends to recognize the fundamental principles and processes of the natural world.

TECHNOLOGY
The integration of scientific knowledge for human needs is known as technology. This integration provides a pathway to the development of system, techniques, and tools.

ENGINEERING
Engineering is the application of scientific principles to construct and improve system, structures, and machine.

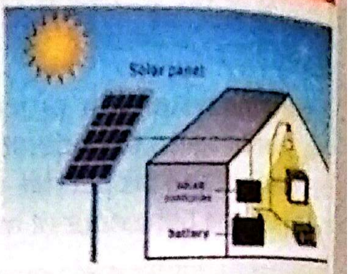
1.6: APPLICATIONS OF SCIENCE, TECHNOLOGY AND ENGINEERING

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 large scale. Together, science, technology, and engineering contribute to the development and utilization of sustainable energy sources.

Example 1.3: 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 1.4: 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 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 t 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 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) Differentiate between science and technology.

(ii) Differentiate between geochemistry and astrochemistry.

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

(iv) How does forensic chemistry help us.

(v) What are the uses of nuclear chemistry?

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 traits. 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 STATES OF MATTER

A state of matter refers to the distinct physical forms that matter can take, characterized by different physical properties like shape, volume, and the arrangement of particles. The four commonly known states of matter are solid, liquid, gas, and plasma. However, there are also more exotic or intermediate states that arise under specific conditions.

1. **Solid:** A solid has a definite shape and volume. Particles are tightly packed in a fixed arrangement. For example, Ice, diamond, gray fight , metals etc .
2. **Liquid:** A liquid has a definite volume but takes the shape of its container. Particles are close together but can move past one another. For example, water, oils, milk, honey etc.
3. **Gas:** A gas has neither a definite shape nor volume. Particles are far apart and move freely. For example, Oxygen gas (O_2), hydrogen (H_2), nitrogen(N_2) etc.
4. **Plasma:** A highly ionized gas where electrons are separated from atoms, often found in high-energy environments. Plasmas are electrically conductive. For example, the sun, lightning, welding arc, solar, wind etc.

Exotic or Intermediate States:

1. **Bose-Einstein Condensate (BEC):** A state of matter that occurs at near absolute zero temperatures, where particles occupy the same quantum state, acting as a single quantum entity. For example, BEC of rubidium atoms at extremely low temperatures (-273.14°C).

   Liquid Crystals: Intermediate states between liquids and solids, where molecules flow like a liquid but retain some ordered structure, commonly used in display technology. For example, Liquid crystal displays (LCDs).

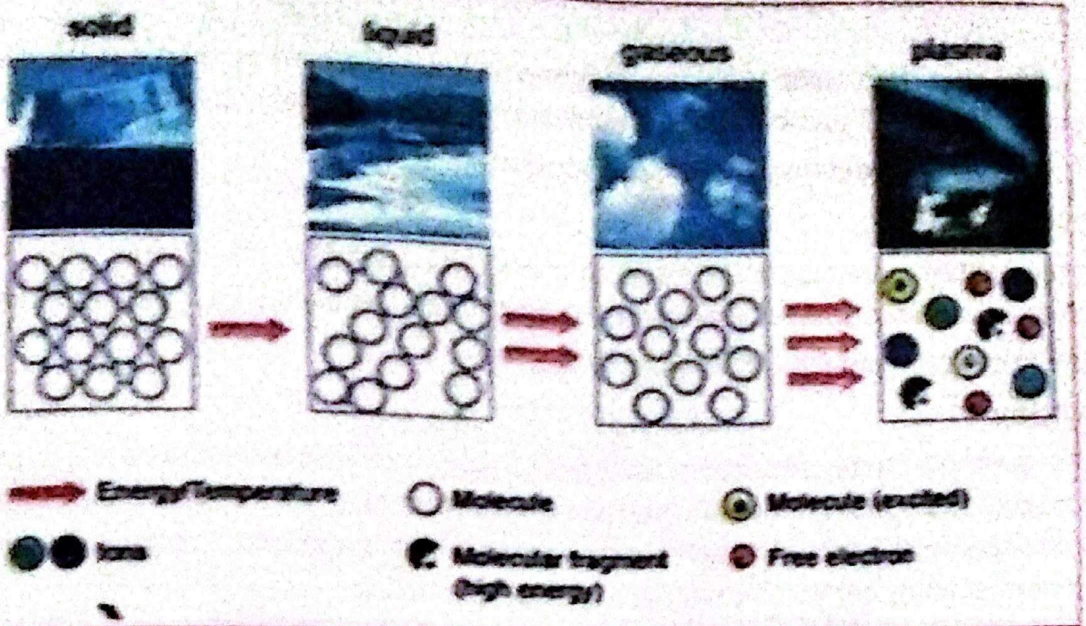


Fig. 2.1: States of Matter

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

Table 2.1: Properties of different states of matter

	States of matter		
Physical Property	**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. A Pure substance consists of single type of particle with a fixed composition and distinct chemical properties.

Matter is classified into elements, compounds and mixtures:

1. **Element:**

   An element is a pure substance made up of only one type of atom. It cannot be broken down into simpler substances by chemical means. Elements contain only one type of atom.

   For examples, Oxygen (O_2), Gold (Au), Carbon (C).

2. **Compound:**

   A compound is a pure substance made up of two or more elements chemically combined in a fixed ratio. Compounds have different properties from the elements that make them and can only be separated into their elements through chemical reactions. Compounds consist of two or more elements chemically bonded in fixed proportions.

   For examples, Water (H_2O), Carbon dioxide (CO_2), Sodium chloride (NaCl).

3. **Mixture:**

   A combination of two or more substances (elements or compounds) that are physically mixed together but not chemically combined is called as a mixture. Mixtures can be separated into their components by physical means, and they do not have a fixed ratio. Mixtures consist of multiple substances that retain their individual properties and can be separated by physical processes

   Examples: Air (a mixture of nitrogen, oxygen, carbon dioxide), Saltwater (a mixture of salt and water), Sand and iron filings.

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.

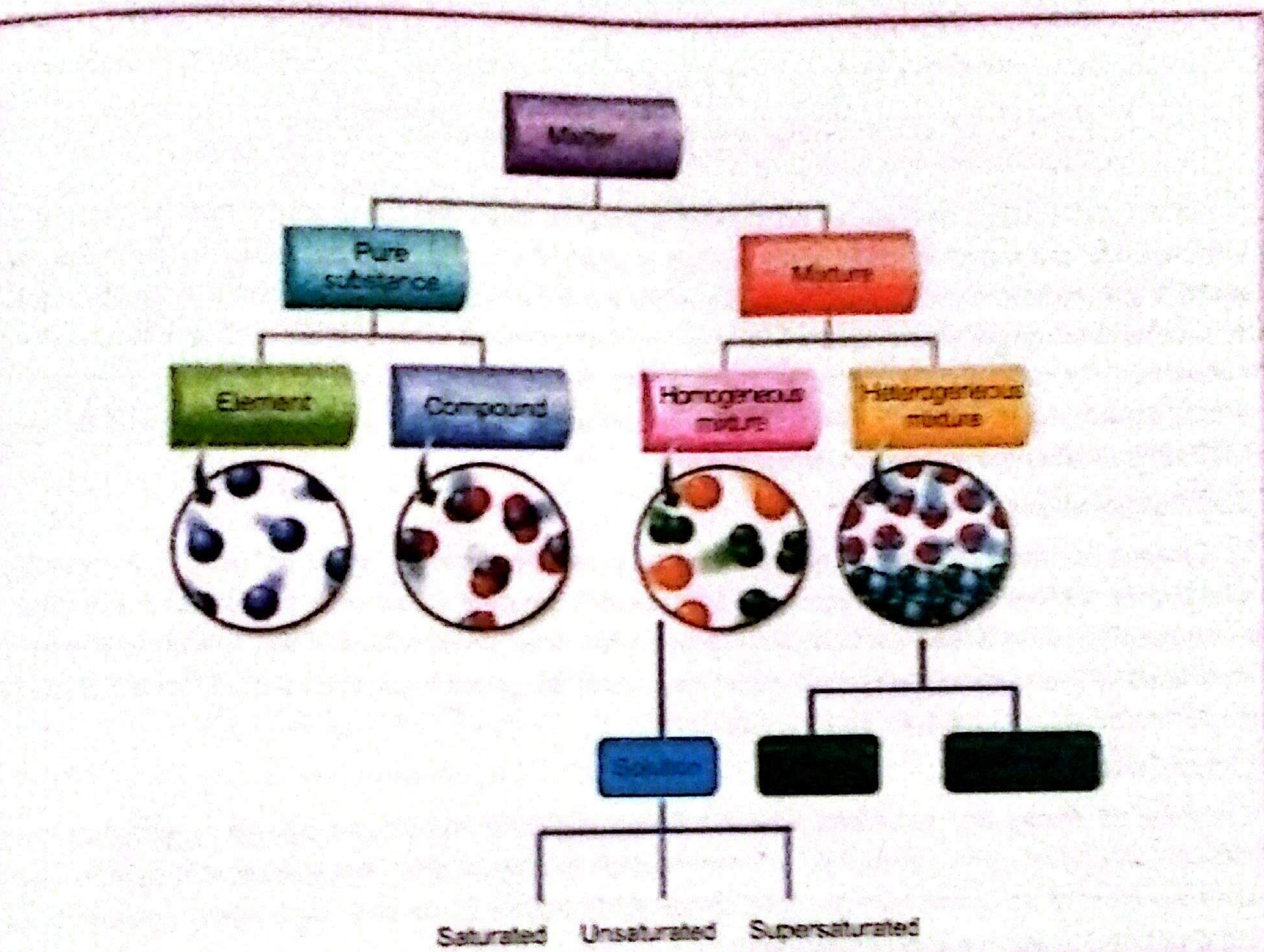


Fig. 2.2: Classification of matter

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.

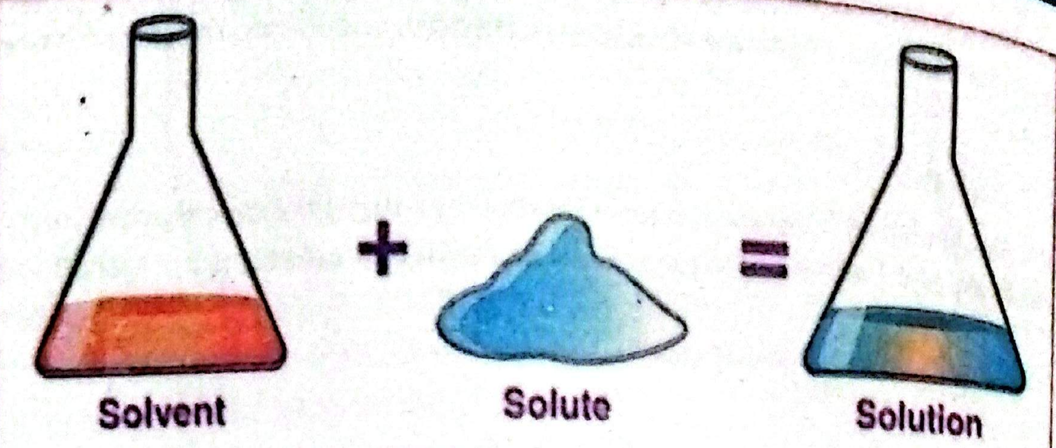


Fig. 2.3: Formation of Solution

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 respectively.

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 Solution

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 (figure 2.4), crystallization will start. When crystallization has finished, you will have a saturated solution in the presence of sodium thiosulphate crystals.

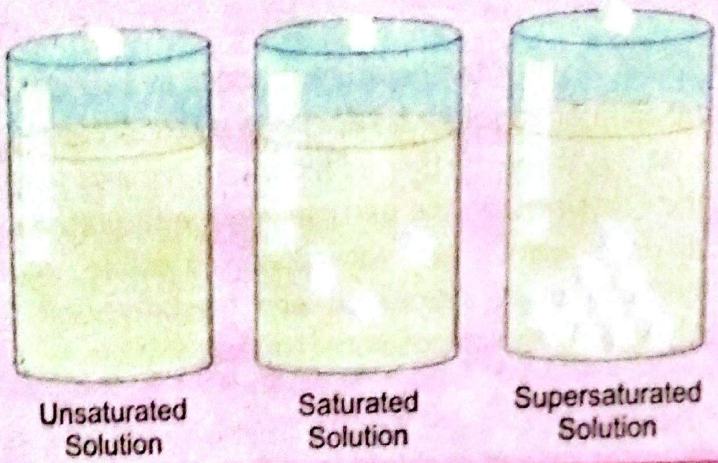


Fig 2.4: Different types of solutions

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

Example:

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

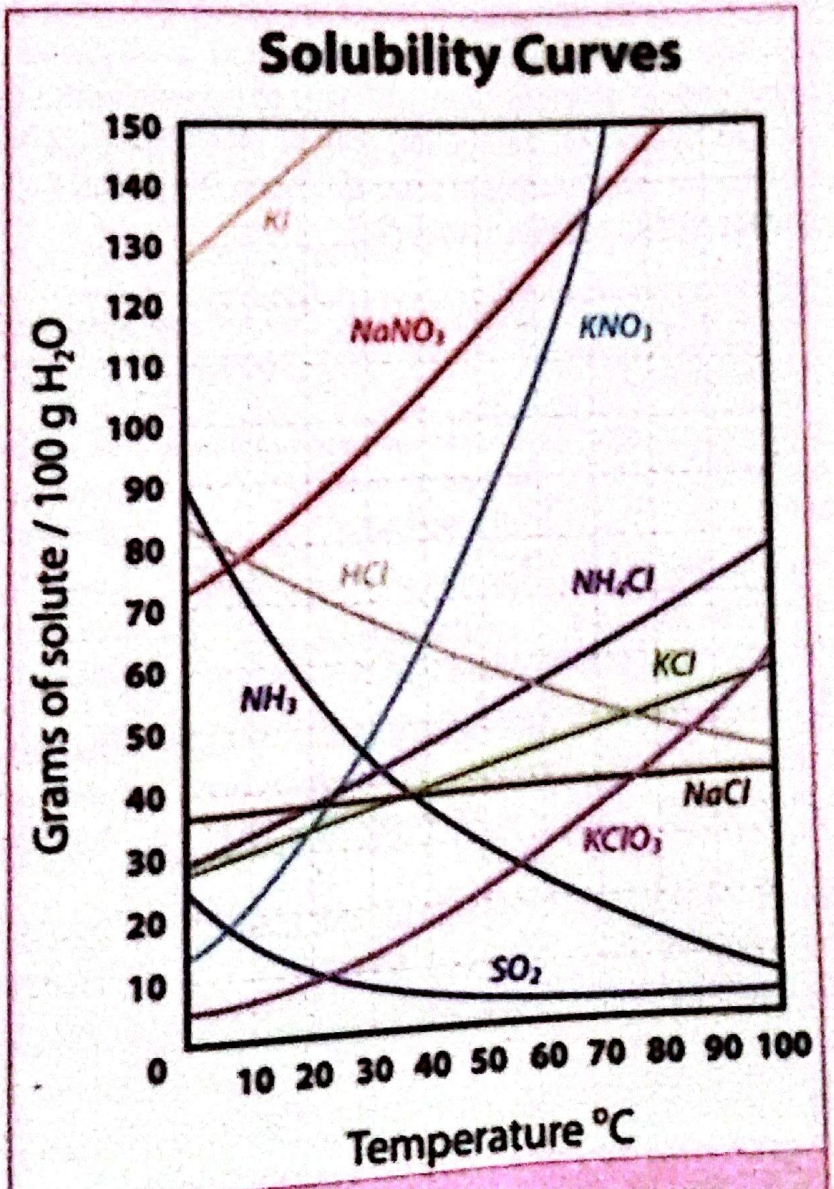


Fig.2.4 Effect of temperature on solubility

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 1000 nm) 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: A Comparison of Solution, Colloid and Suspension

S.No	Solution	Colloid	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 is less than 1nm. Not visible by naked eye	Particle size vary from $1\text{-}10^3$ nm Visible by naked eye	Particle size greater than 10^3 nm. Visible 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 (Tyndall effect)	Can Scatter the light (Tyndall 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 elements 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 allotrope. Explain the allotropes of carbon.
4. What is the difference between homogeneous and heterogeneous solution?
5. Differentiate between the colloids, suspension and solution.
6. If there are 18 protons in the argon atom, then what is the atomic number of argon?
7. Describe states of matter with examples.
8. Differentiate between the following:
   a. Colloids and Suspensions
   b. Elements and Compounds
   c. Concentrated and Dilute solutions
9. Examine the concept of solubility.

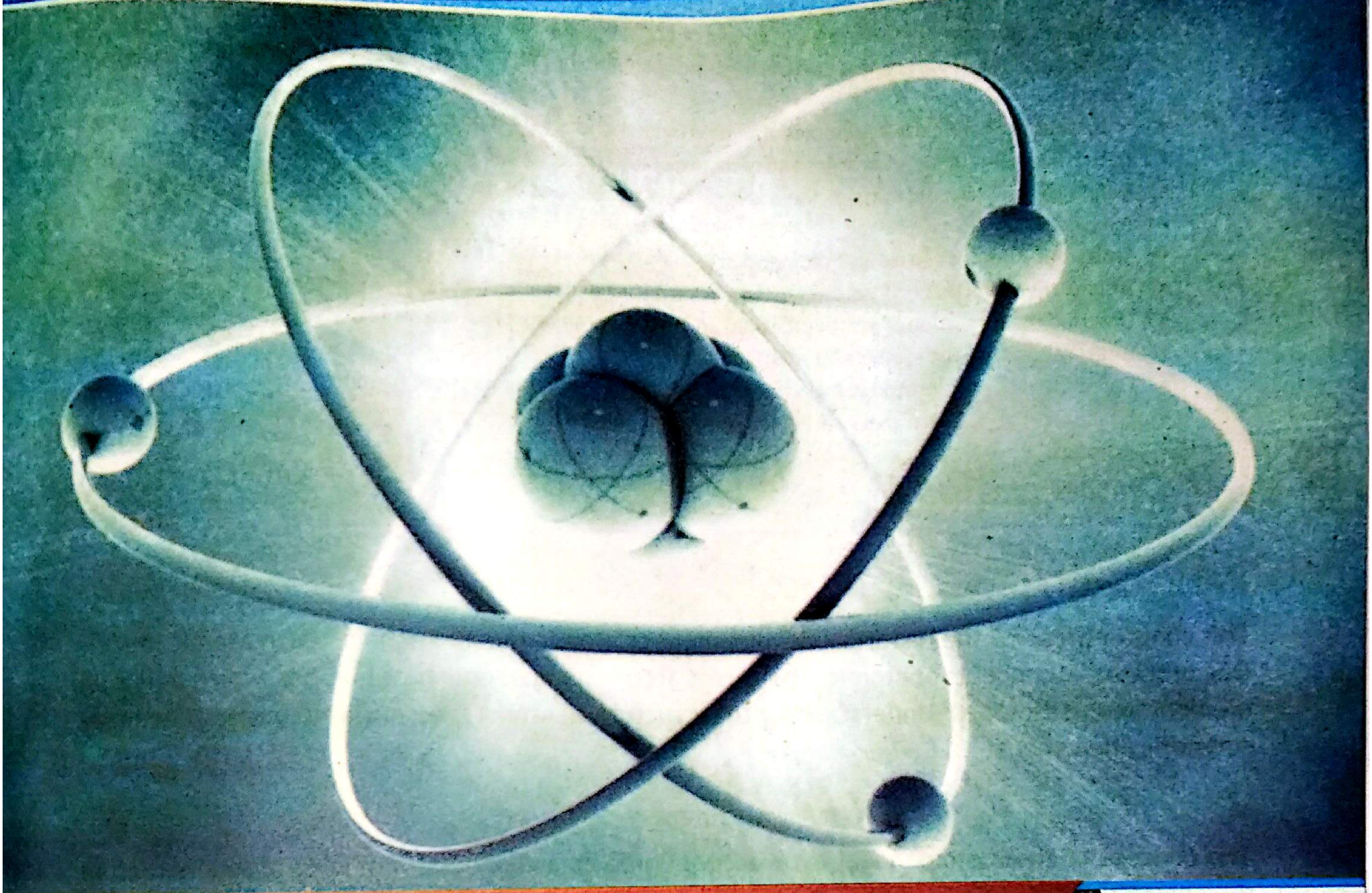
THINK TANK

10. **Why is a solution considered mixture?**
11. **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 understand 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 rearrange. They combine in simple ratios.
4. Atoms can neither be created nor be 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 0.00004 cm thickness with α-particles. (fig. 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. Around 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.

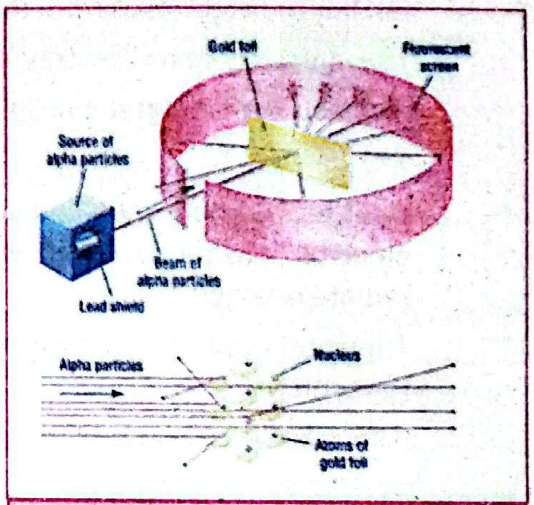


Fig. 3.1 A three-dimensinal view of an apparatus clearly showing scattering of alpha particles by gold

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.

However it also explains the observed line spectrum of the hydrogen atom. Main postulate Bohr's atomic theory are as follows:

1. The electron in an atom revolves around the nucleus in one of the circular orbits. E 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. 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 electron is an integral multiple of $\frac{h}{2\pi}$ where h is Plank's constant (its value is 6.626x1 J.s).
4. Light is absorbed when an electron jumps to a higher energy orbit and emitted when electron falls into a lower energy orbit. Electron present in a particular orbit does n radiate energy.
5. The energy of the light emitted is exactly equal to the difference between the energies 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

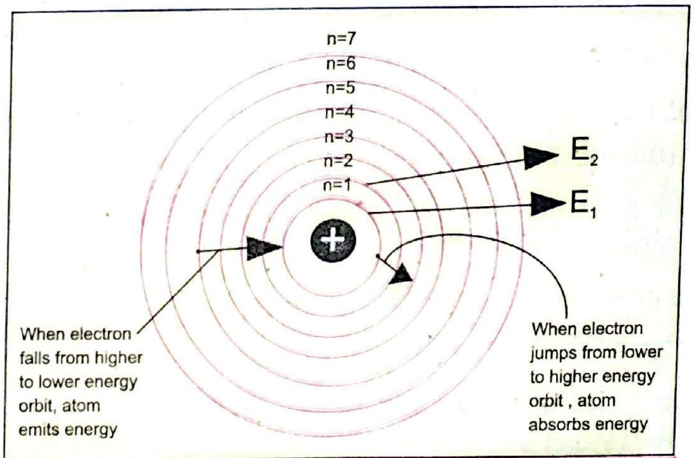


Fig. 3.2: Bohr's model of the atom

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. incorporates the principles of quantum mechanics and treats electrons as wave-particle entitie Instead of exact orbits, it defines probability regions, called orbitals, where electrons are like to be found.

The Heisenberg Uncertainty Principle

Heisenberg uncertainty principle is one of the fundamental concepts of quantum mechanics a 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 a future trajectory of an electron. As a result, plotting the electron orbit around the nucle 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 ma 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? fig 3.3

- 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.

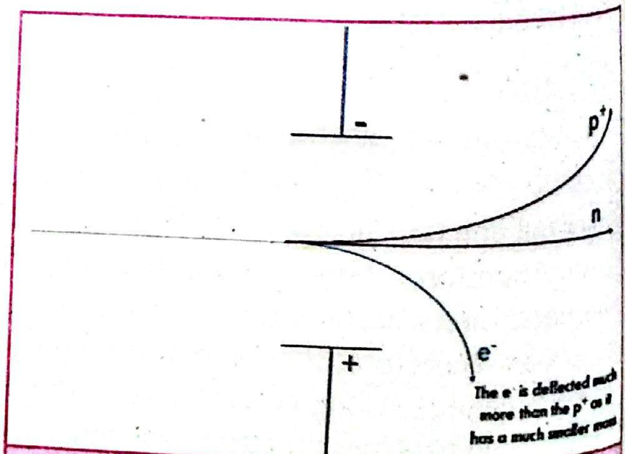


Fig. 3.3: Path of positively, negatively charged particles and neutral particles through the uniform electric field.

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

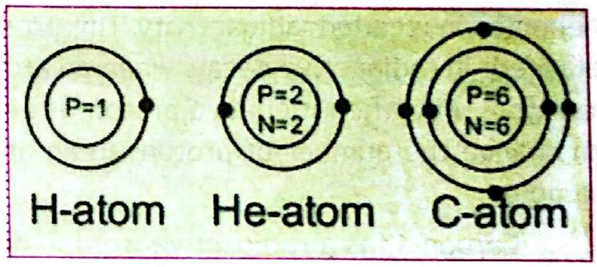
Different isotopes of the same element have the same number of protons in their atomic nuclei but differing numbers of neutrons. Some isotopes of an element are unstable and show radioactive decay. Radioactive isotopes of an element can be defined as atoms that contain an unstable combination of neutrons 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 the 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

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.

Mass of one C-12 atom = 12 amu

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

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

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

$$= 1.008\ amu$$

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

3.5 ISOTOPES

Fig. 3.1 shows Bohr's Model for two atoms of Ne. Can you identify three similarities and two differences in these atoms?

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.

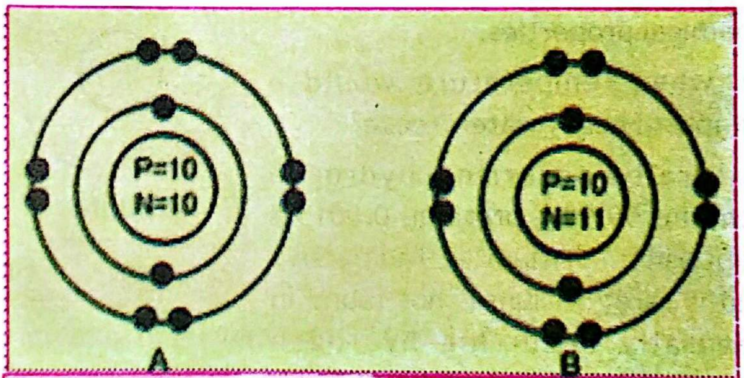


Fig. 3.1: Two isotopes of Ne

(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 Physical Properties of ordinary water and heavy water

Physical property	Ordinary water	Heavy water
Melting Point	0.00 °C	3.81°C
Bioling point	100 °C	101.2°C
Density at 25°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.

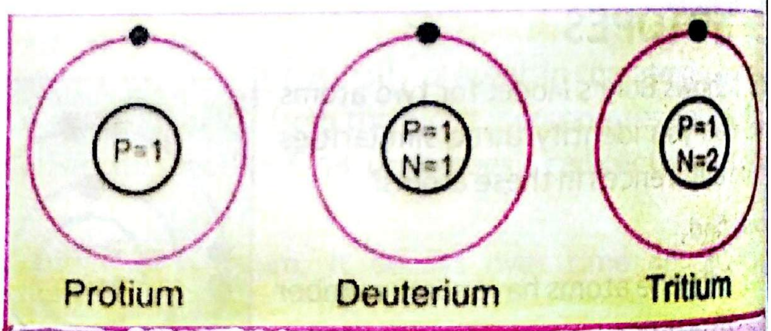


Fig. 3.2: Isotopes of hydrogen

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?

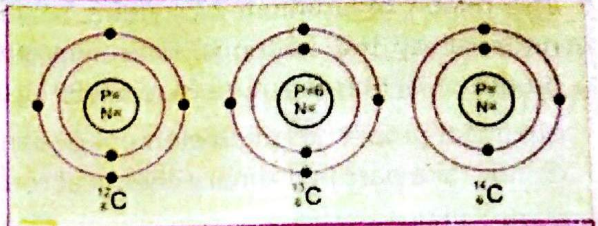


Fig. 3.3: Isotopes of Carbon

Natural abundance of isotopes of carbon is as follows:

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

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%

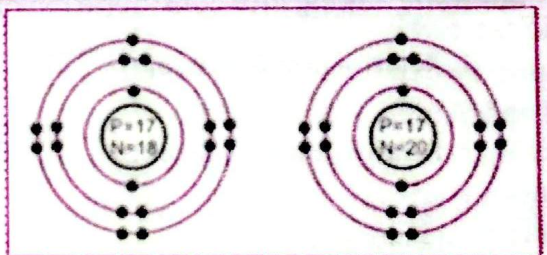


Fig. 3.4: Isotopes of chlorine

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 ene[rgy] released by the fission of one kilogram of uranium-235 is equivalent to chemical ene[rgy] produced by burning more than 17000 kg of coal.

Chemical properties of an element depend upon the number of protons and electrons. Neutro[ns] do not take part in ordinary chemical reactions. Therefore, isotopes of an element have simila[r] 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 it[s] isotopes and their natural abundance.

Example 3.2: 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 a[n] element, taking into consideration of their natural abundance. Use general formula

$$\text{Relative atomic mass of C} = \frac{\text{NA of C-12} \times \text{A of C-12} + \text{NA of C-13} \times \text{A of C-13} + \text{NA of C-14} \times \text{A of C-14}}{100}$$

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

NA= Natural abundan[ce]
A= Atomic mass

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

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

CONCEPT ASSESSMENT EXERCISE 3.1

An element has two isotopes A and B.
The relative atomic mass of element is 35.5 amu. Natural abundance of isotope A is 75.77 % and its isotopic mass is 35. Find the isotopic mass of B if its natural 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 3.3: Describing the formation of cations

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

Problem Solving Strategy:

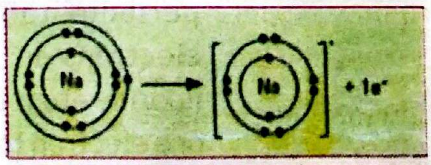
1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. Sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group IIA in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

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

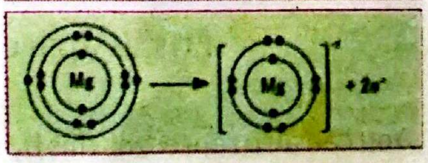
You can also represent this by following electron dot structure,



(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 3.4: 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.

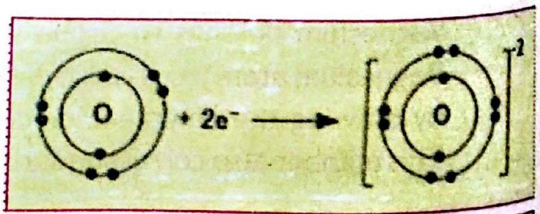
Solution:

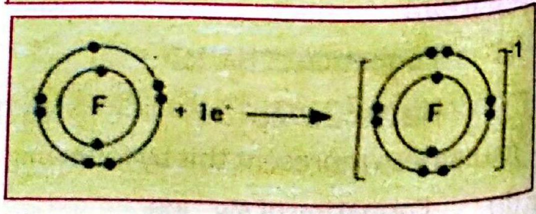
(a) Formation of anion by oxygen atom.

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.

There are total 7 shells
1. K 2. L 3. M 4. N 5. O
6. P 7. Q

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 .....

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.

Shell	Number of Subshell
K	1s
L	2s, 2p
M	3s, 3p, 3d
N	4s, 4p, 4f

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) $1s^2, 2s^2, 2p^7$
(c) $1s^2, 2s^2, 2p^5, 3s^2$ (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. ✓done on NIB.
(ii) Why an atom is electrically neutral? Because No. of protons = Electrons.
(iii) How many sub-shells are there in N shell. 4s, 4p, 4d, 4f
(iv) Give notation for sub-shells of M shell. 3s, 3p, 3d
(v) List the sub-shells of M Shell in order of increasing energy. 3s < 3p < 3d
(vi) Can you identify an atom without knowing number of neutrons in it? on NIB.

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

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 → 2s (b) 3p, 2p → 2p (c) 3s, 4s → 3s

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

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~~) 40

9. Write electronic configuration for the following elements: on N/B

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

11. State the importance and uses of isotopes in various fields of life. on book pg 38, 39

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

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? Down ↓

Q12: Solution,

Atomic No = 23

Mass No = 56

Atomic Number = Number of protons.

Q13:

Atomic Symbol Shows the atomic Mass / Mass No and atomic number.

Atomic No = 13

Atomic Mass / Mass No = 27

So, Protons = 23

In neutral atom protons = electrons

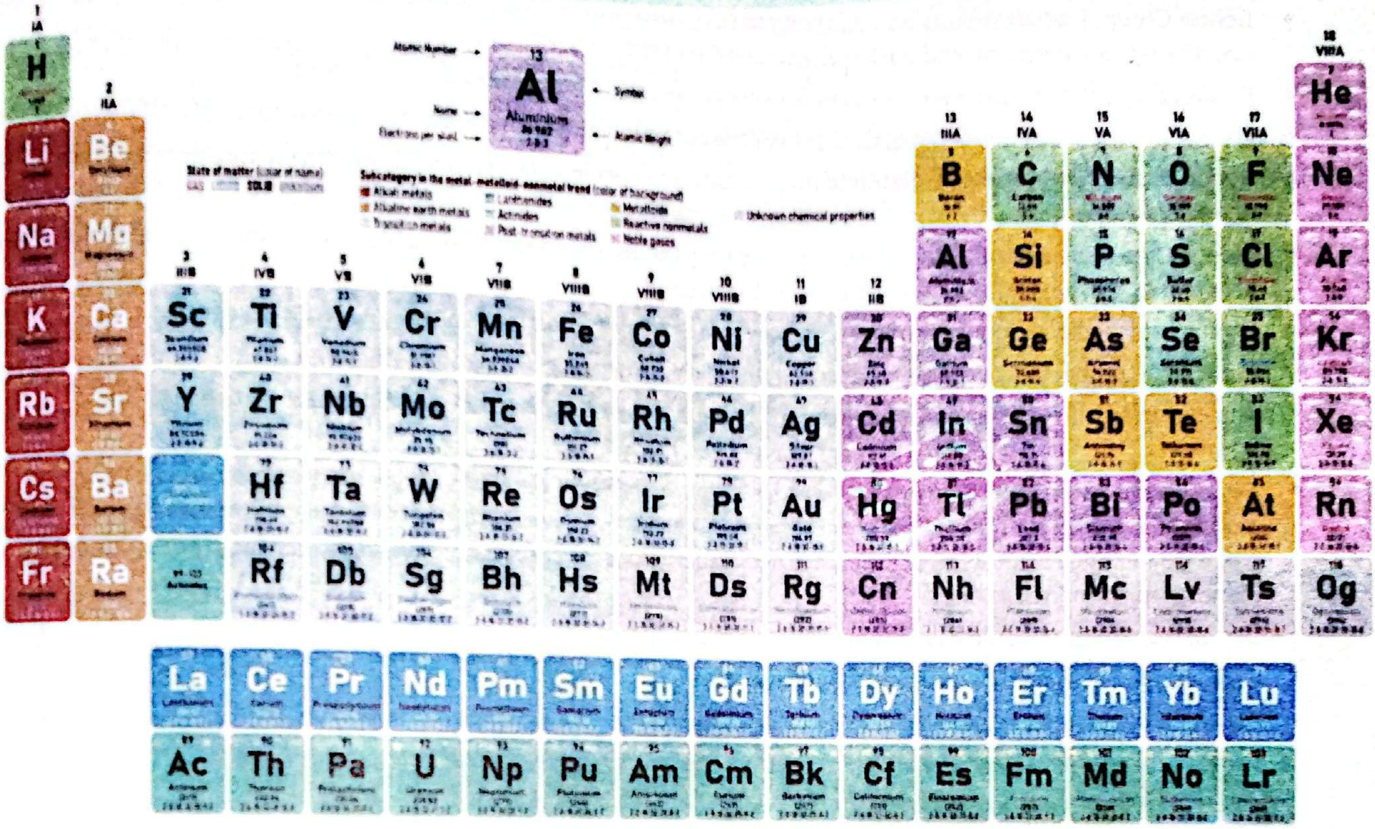
Hence Electrons = 23

Neutrons = A − Z (Atomic Mass − Atomic No)

= 56 − 23

= 33

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 and 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

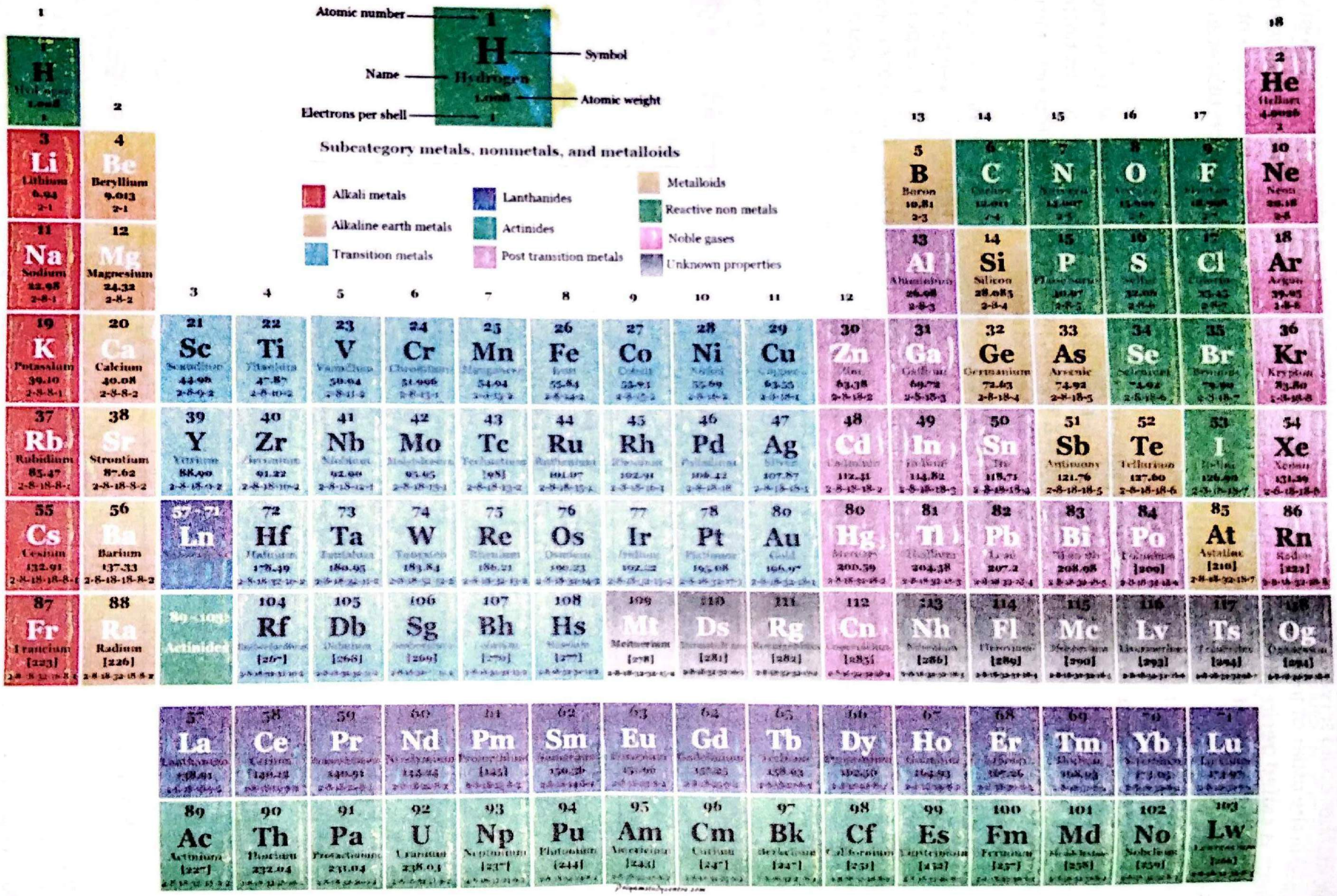
Activity 4.1

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	

Fig. 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.**

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 are called halogens. The Group 18 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 3 period.

Since Na belongs to the s block, its group number is equal to a number of electrons in valence subshell. 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. Electronic 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 sub-shell 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-shell.

   = 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 number.
3. The elements whose atoms have same value of n for the valence shell lie in the same period.

Solution:

	2	15	17	
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 of the following table.

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

$Q = 1s^2, 2s^2, 3p^1$

$R = 1s^2$

$S = 1s^2, 2s^2$

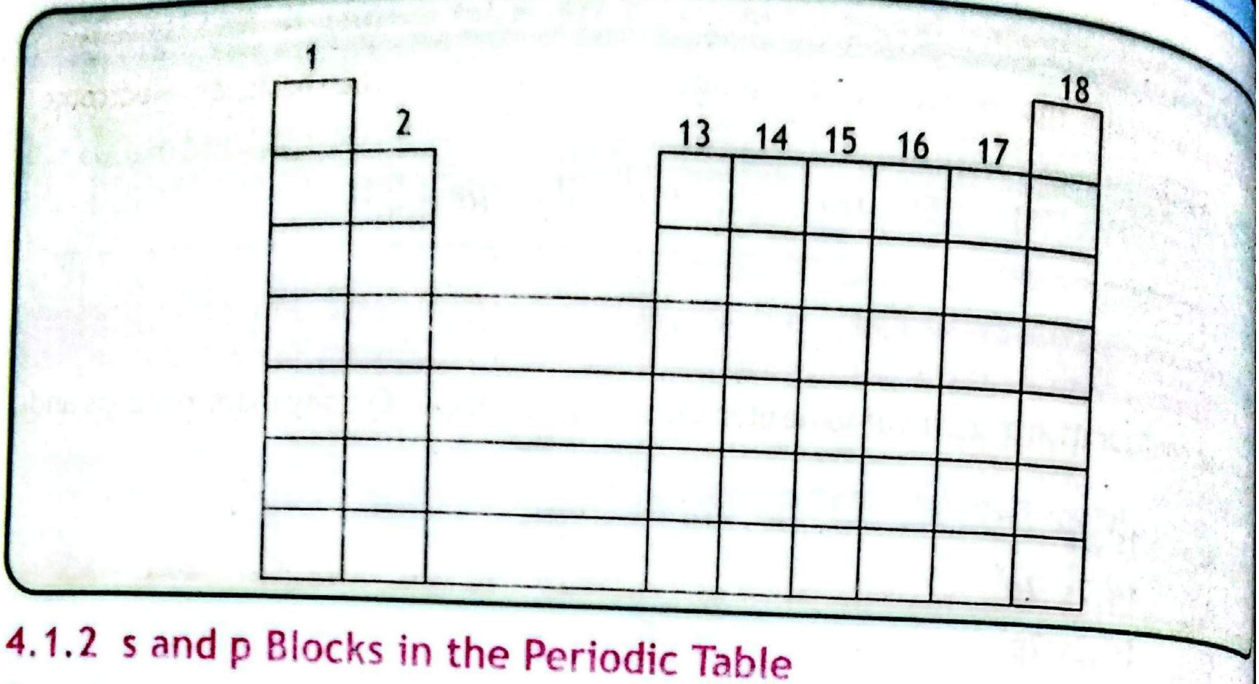
$T = 1s^2, 2s^1$

$W = 1s^2, 2s^2, 2p^6$

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

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

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



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 f-orbital is in the process of completion. Figure 4.2 shows the blocks of the periodic table.

Li = $1s^2, 2s^1$, 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-blocks, 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^-.

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

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

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

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

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 (Z=7) (b) Oxygen (Z=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

∴ 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

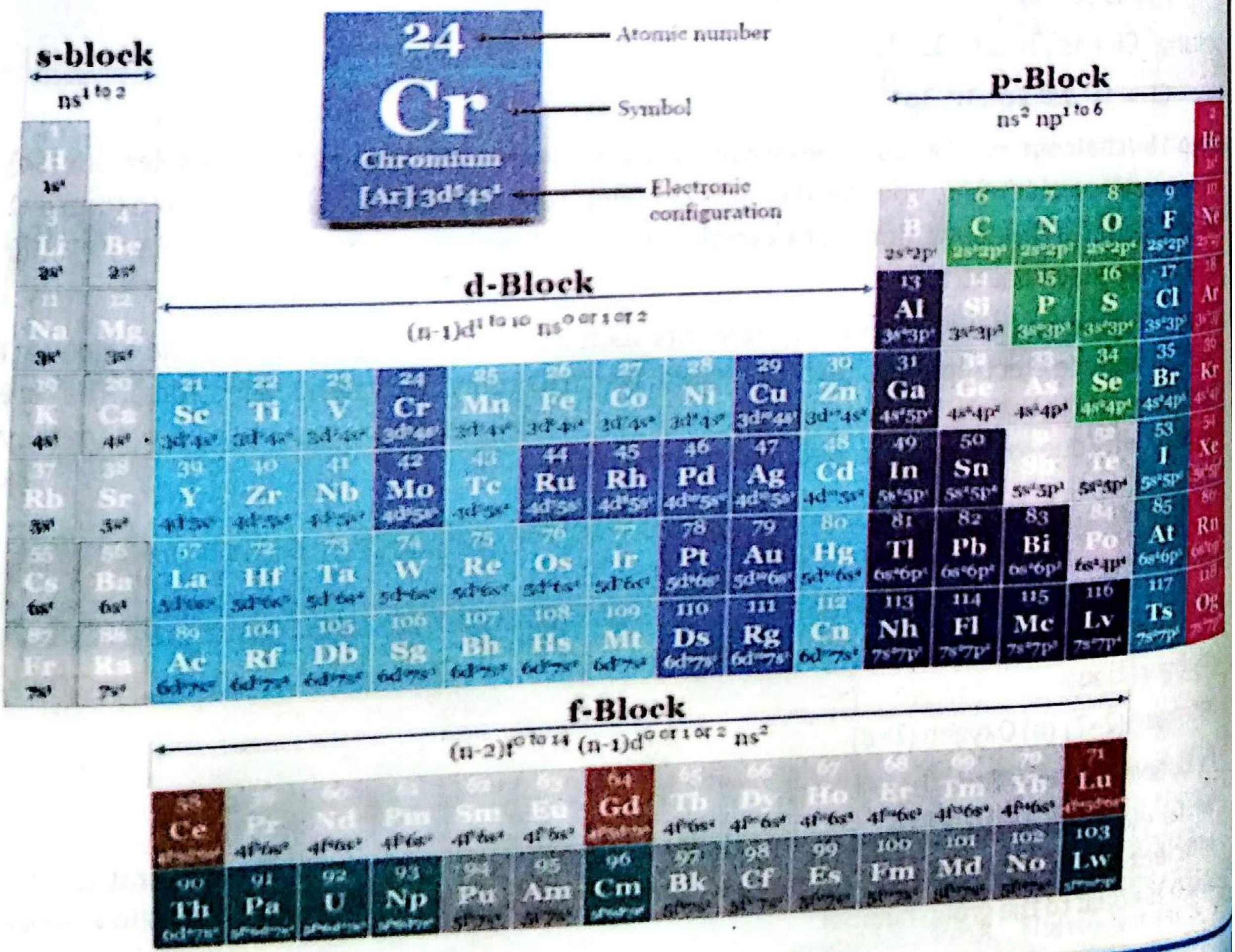
Number of valence electrons = 2 + 4 = 6

O belongs to p-block

So, Group number = 6+10 = 16

Oxygen is present in the 2[nd] period of Group 16

Fig. 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.

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.

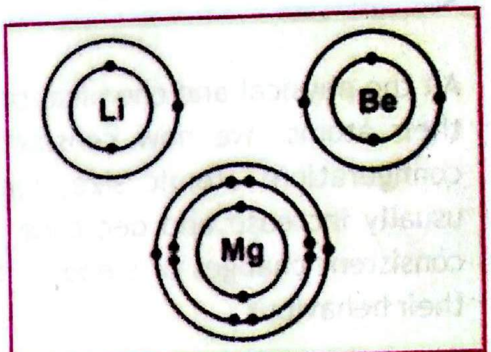


Fig. 4.3: Electronic structure of Li, Be and Mg

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 their behaviour.

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

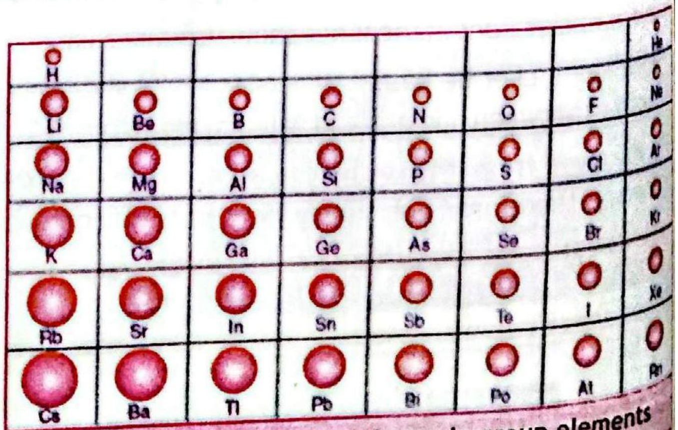


Fig. 4.4: Atomic sizes of the main group elements

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".**

$$M_{(g)} + \text{ionization energy} \longrightarrow M^{+}_{(g)} + \bar{e}$$

Period No.	1	2	13	14	15	16	17	18
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

Fig. 4.5: Ionization energy values 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.

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. However little by little the nuclear charge increases and atomic size decreases. The stronger attraction between the nucleus and the valence electron increases. As a result, the ionization energy generally 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

Fig. 4.6: Electron affinity values of main group elements

There are several exceptions to the general trend of election affinity values. You wil 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 Figure 4.7 shows as a scale of electronegativities of the elements devised by Linus Pauling American chemist Linus Pauling devised a method for calculating the relative electronegati 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						

Fig. 4.7: The Electronegativity values of main group 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 1 and 2 and note down the variation in electronegativites value.

5. Move from top to bottom in Groups 16 and 17 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 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 hydroxide and hydrogen.

  metal + water ---> metal hydroxide + hydrogen gas

- 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 metal hydroxide 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 elements have one electron in valence sub- shell s. As n value of valence sub-shell is 4, the unknown element must lie in the 4th 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 19.3 g/cm^3 at room temperature.

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 °C, platinum(Pt) has a melting point of 1768 °C.

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 and 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 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.

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.

2. 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^-.

3. 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 between molecules 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

4. 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:

$$F_2 > Cl_2 > Br_2 > I_2.$$

5. Displacement reactions of halogens

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

6. Hydrogen halides and their thermal stabilities

Halogens react with hydrogen to form hydrogen halides.

$$H_2 + X_2 \rightarrow HX \qquad \text{Where } X = F_2, Cl_2, Br_2, I_2$$

The strength of the hydrogen-halogen bond is related to the electronegativity differe between the hydrogen and halogen atoms. A larger electronegativity difference results stronger bond. As we move from HF to HI, the electronegativity difference between the hydr and halogen atoms decreases, resulting in weaker bonds in HCl, HBr, and HI. So, the rela 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 17

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

Activity 4.6

Predict the properties of other elements in Group 17, from the given data given above.

From this data you can predict how the halogens will behave up and down the group. Astat the fifth halogen, will have high melting and boiling points, so, will be solid at r 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 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	Electron Configuration
Helium	2	$1s^2$
Neon	10	$1s^2, 2s^2,2p^6$
Argon	18	$1s^2, 2s^2, sp^6,3s^2, 3p^6$
Krypton	36	$1s^2,2s^2, 2p^6, 3s^2, 3p^6, 4s^2,3d^{10},4p^6$
Xenon	54	$1s^2,2s^2, 2p^6,3s^2,3p^6,4s^2,3d^{10}, 4p^6,4d^{10},5s^2,5p^6$

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

Here is a comparision of general properties of metals and non-metals.

1. 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.

2. 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.

3. 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 slip pass each others easily under pressure. Non-metals are neither malleable nor ductile rather they are brittle.

4. 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 1 and 2 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 13 to 18 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) 1 (b) 2
(c) 17 (d) 18

(iii) Which of the following elements belongs to 18?

(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 2 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 13.
   (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

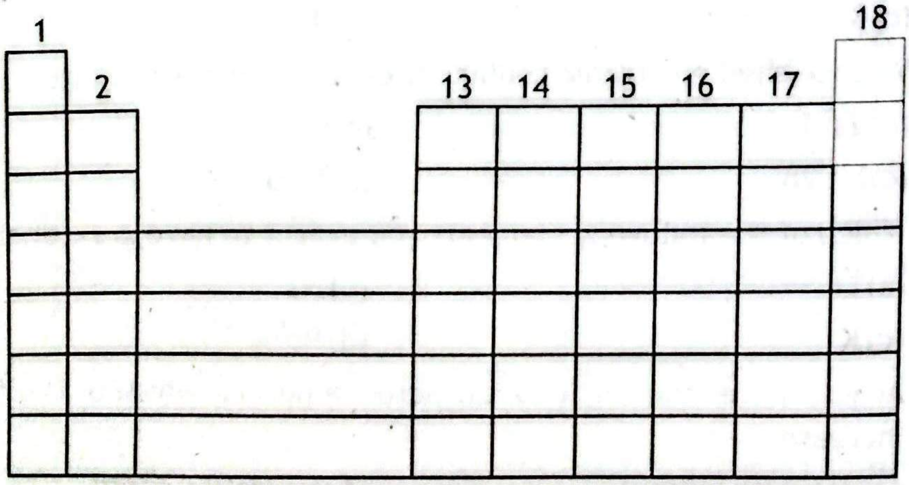
5. Specify which of the following elements you would expect to have the greatest electron affinity.

   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 15

(b) An element present in period 2 of Group 16

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) $1s^2$

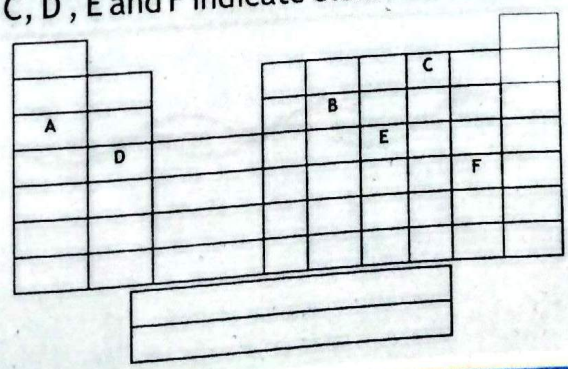
THINK TANK

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 configurations. Give reason for your choice.

ii. Use the second member of each group from Group 1, 2 and 17 to judge that the number of valence electron in an atom of the element is the same as its group number.

iii. Letters A, B, C, D, E and F indicate 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 that in what region of the periodic table you will find elements with relatively:

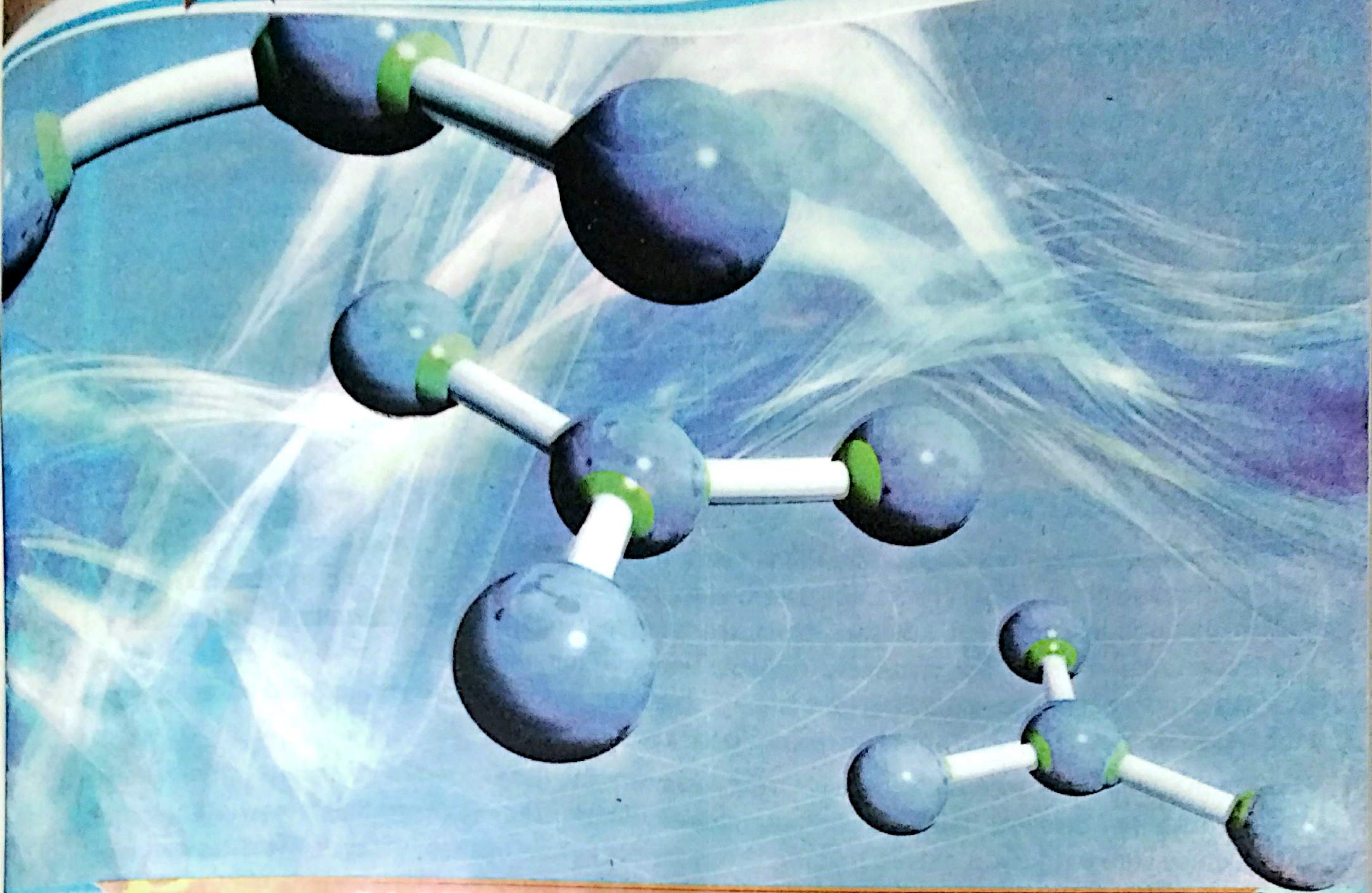
a) high ionization energies

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 (1, 2, 13, 14, 15, 16 17 18) in the periodic table. Group 18 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.

$_{2}He = 1s^2$

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

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

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 halogens follow the octet principle. Hydrogen, helium, and lithium follow the duplet rule because their electrons lie in s orbital.

$_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$ (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) $_{17}Cl^- = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ 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:

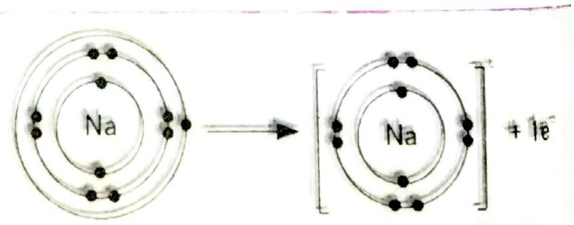
1. Sodium belongs to Group 1 in the periodic table. It has only one electron in its valence shell. The sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group 2 in the periodic table. It has two valence electrons. The 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

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

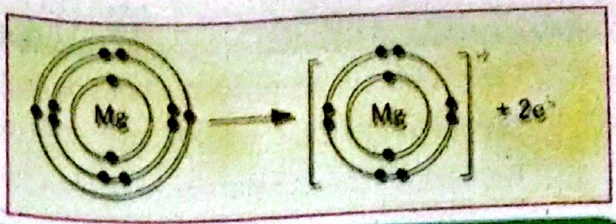
You can also represent this by following electron dot structure,



(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(Z = 3)
   (b) Al(Z = 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(Z = 8) (b) Fluorine (Z = 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.

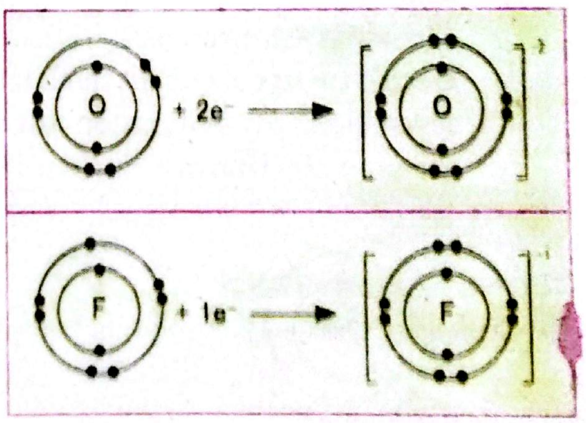
Solution:

(a) Formation of anion by oxygen atom.

Oxygen belongs to Group 16 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,



(b) Formation of anion by fluorine atom

Fluorine belongs to Group 17 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

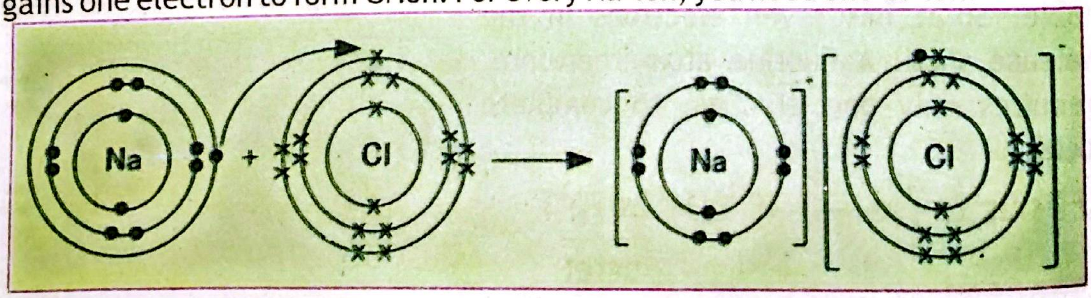
Problem Solving Strategy:

1. The metal atoms form cations and non-metal atoms form anions.
2. The number of electrons lost by metal atoms of group 1 and 2 are equals to their respective group number and in group 13 the number of electrons is 13-10= 3.
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.

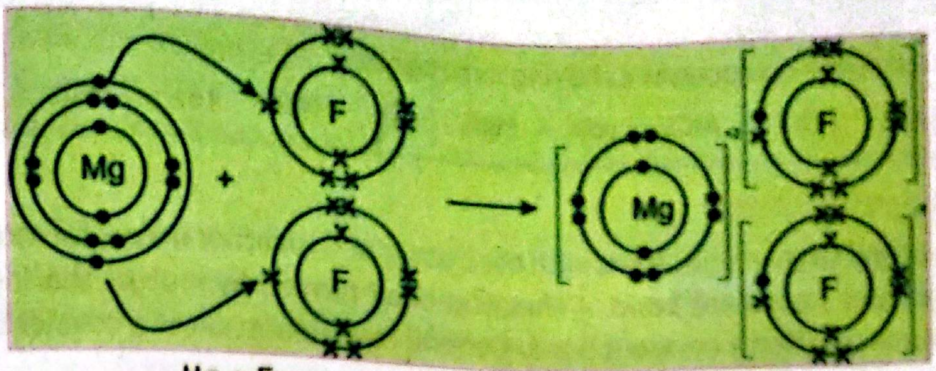
Metal atom tends to lose electrons and non - metal atoms tends to gain electrons to acquire electronic configuration of their nearest noble gas. Since, Na atom has one electron in its outer most shell. It losses one electron to form Na^+ ion. Since, Cl atom has seven electrons in its outermost shell, 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.



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

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

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



$Mg + F_2 \longrightarrow MgF_2$ (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 1 and 2 are equal to their respective 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 its 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 its outmost shell. So it loses one electron to form Na^+ ion. Since a F atom has seven electrons in its 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.

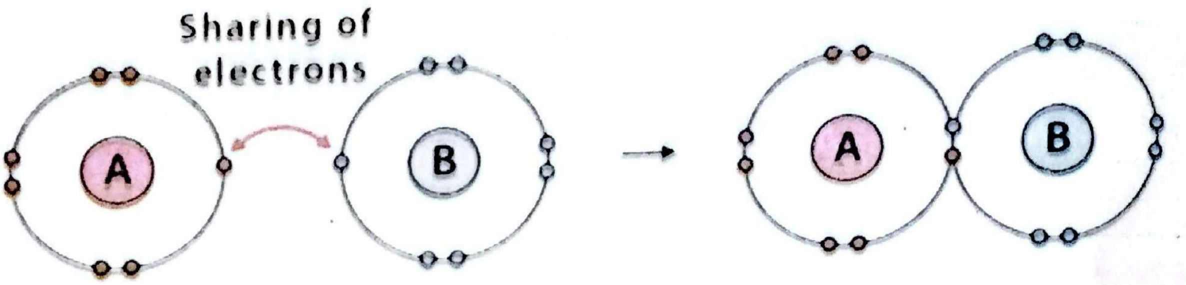
CONCEPT ASSESSMENT EXERCISE 5.4

Recognize the following compounds as having ionic bonds:

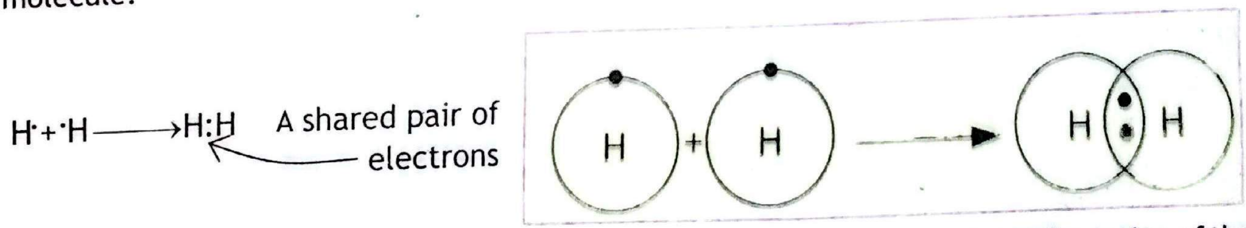
(a) KCl (b) $AlCl_3$ (c) MgF_2 (d) NaF (e) NaBr

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.



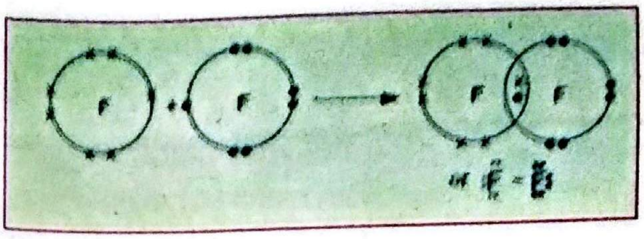
Consider the formation of a covalent bond between two hydrogen atoms. A hydrogen atom has one valence electron. Two hydrogen atoms share their valence electrons to form a diatomic molecule.



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 force of attraction between the bonding electrons and two atomic nuclei binds them together.

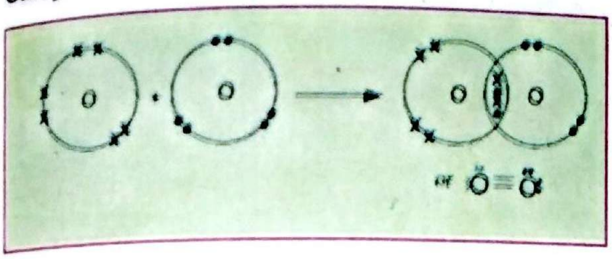
A covalent bond between two atoms can be represented by using electron dot and cross diagram for the atoms and the resulting molecule. As already discussed valence electrons are represented by dots. Just to understand sharing, we represent valence electrons in one atom by dots and in the other atom by crosses. However, remember that all the electrons are identical and cannot be differentiated. A shared pair of electrons is also represented by a solid line in a molecule.



Consider the formation of a bond between two fluorine atoms. Fluorine belongs to Group 17, so it has seven electrons in its valence shell. It needs one more electron to attain the electron configuration of its nearest, Ne 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.

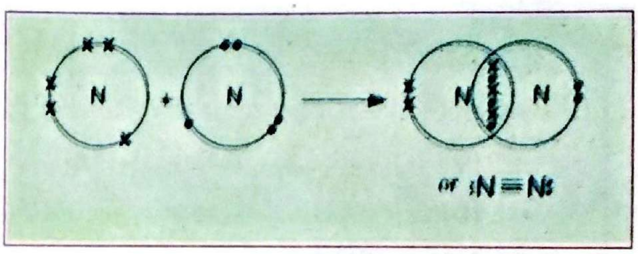
Pairs of valence electrons that are not shared between atoms are called **lone** pairs or unshared pairs. A covalent bond formed by sharing **one pair** of electrons is called a single covalent bond. So both H_2 and F_2 molecules contain **single** covalent bond.

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 16, so it has 6 electrons in its 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 dot and cross diagram for simple covalent molecules containing single covalent bonds

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

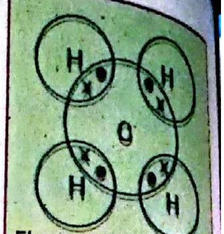
Problem Solving Strategy:

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.

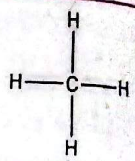
Solution:

(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.



Electron dot and cross diagram of CH_4



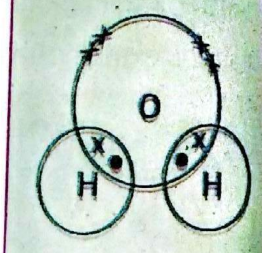
(ii) Connect the atoms with a dot and a cross

(b) H_2O

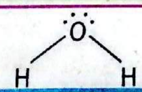
(i) O has six valence electrons $:\ddot{O}\!:$ and each hydrogen atom has one valence electron ($\dot{H}$). So O-atom needs two electrons to complete the octet. Each H needs one electron to complete duplet.

(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)



Electron dot and cross diagram of H_2O



CONCEPT ASSESSMENT EXERCISE 5.5

Draw electron dot and cross diagram for the following molecules:

(a) NH_3

(b) HCl

(c) CH_3OH

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

Draw electron dot and cross diagram 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:

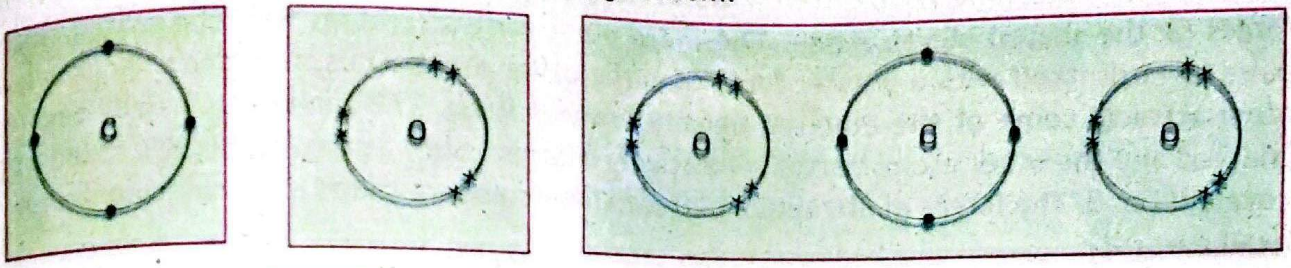
(a) CO_2

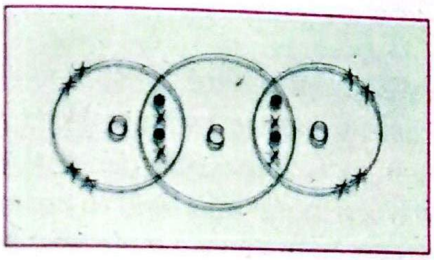
(i) C has four electrons in the valence shell. It needs four electrons to complete octet.

(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.



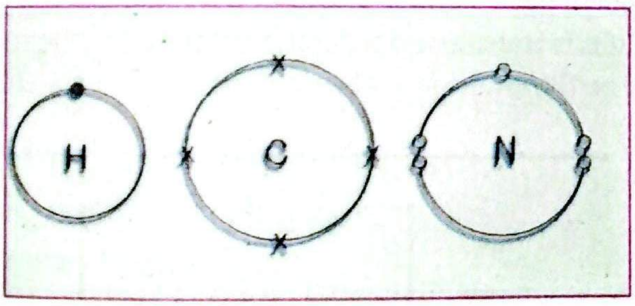


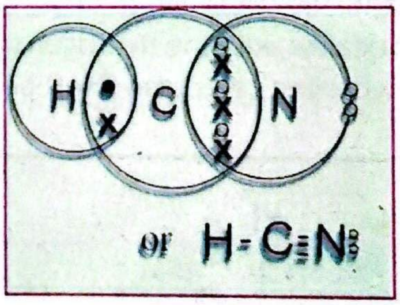
or $:\ddot{O}=C=\ddot{O}:$

(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 dot and cross diagram 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 may **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, H–C̈l:

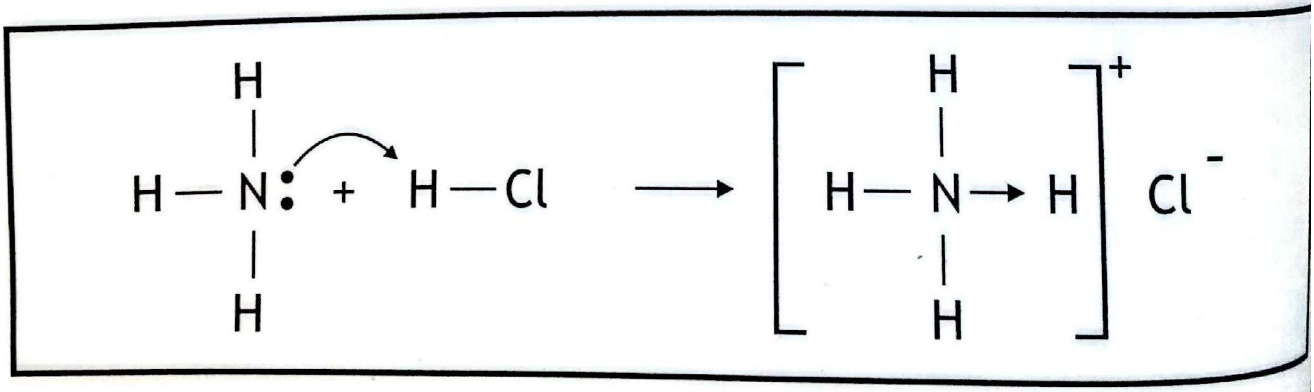
5.3.4 Coordinate Covalent Bond

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 a coordinate 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.

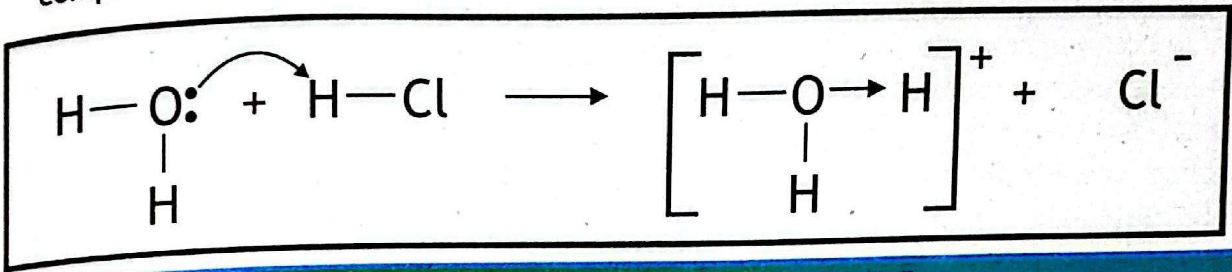


2. **Hydronium ion (H_3O^+)**

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.

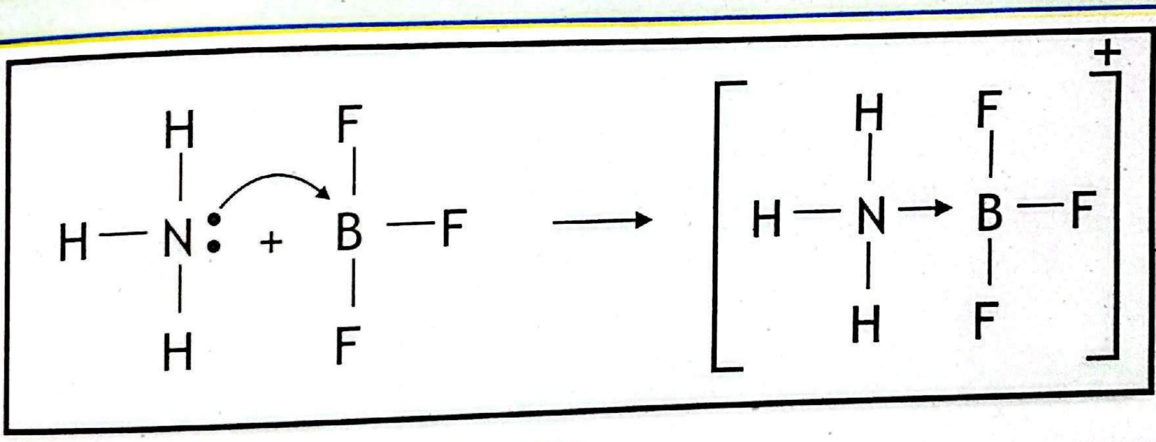
3. **Ammonia Boron Trifluoride (NH_3-BF_3)**

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 complex compound forms through a coordinate covalent bond.



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.

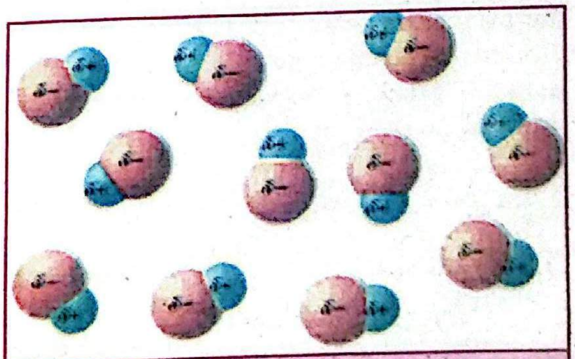


Fig. 5.1: Dipole-Dipole interactions

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 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.

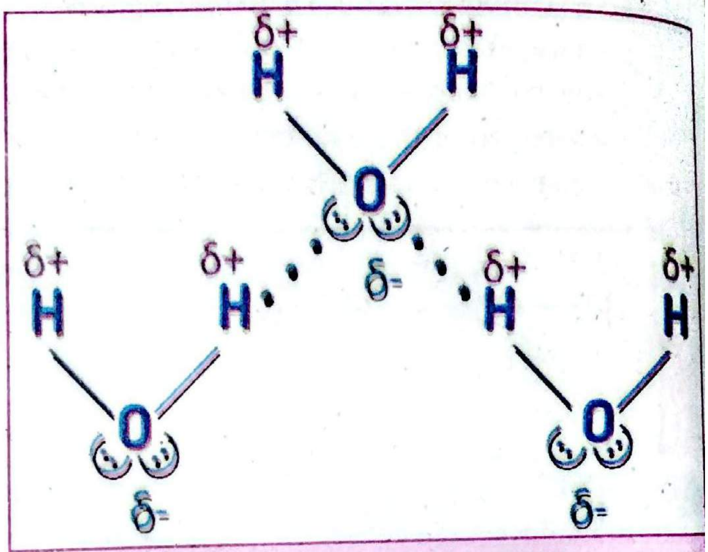


Fig. 5.2: Hydrogen bonding in water

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 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 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. The 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.

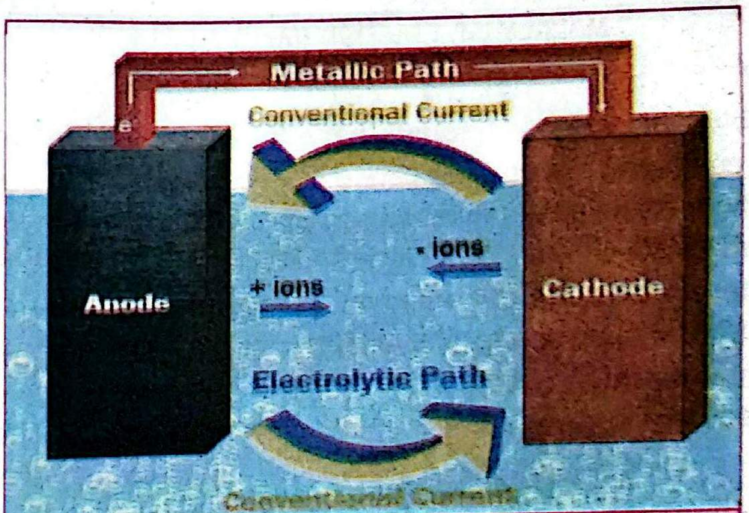


Fig. 5.4: Conduction of electricity through molten NaCl

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 (°C)	Boiling Point (°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 (°C)	Boiling Point (°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 molecules.

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 and electric motors.
5. Graphite material is used in engineering sectors in the making of thrust and journal bearing, piston rings, and valves.
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.

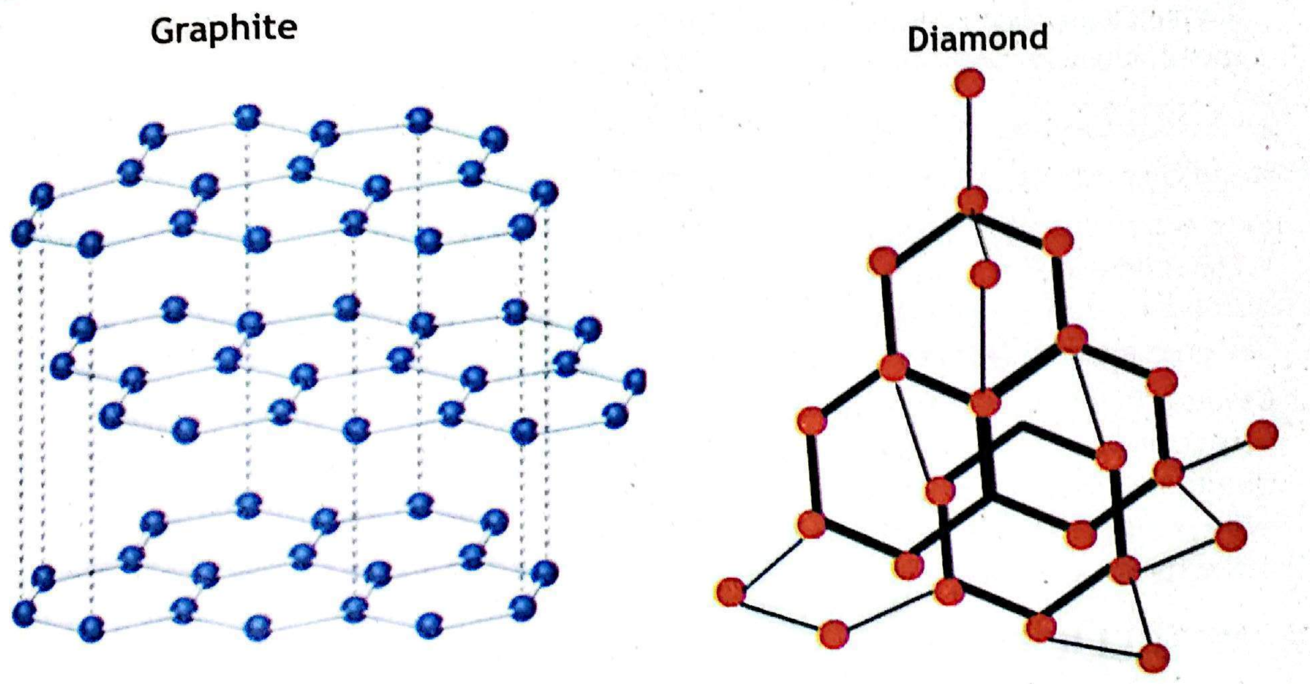


Fig. 5.5: Structure of graphite and dimond

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.

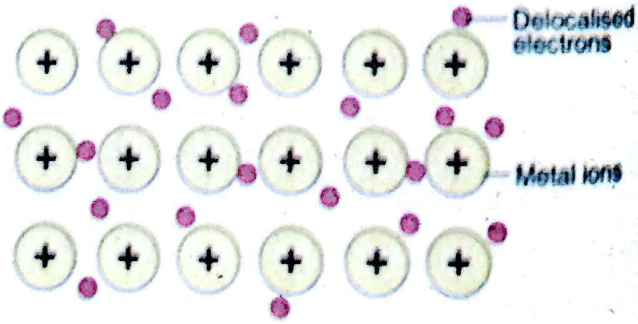
5.5.3 Contrasting ionic and covalent compounds and their uses

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

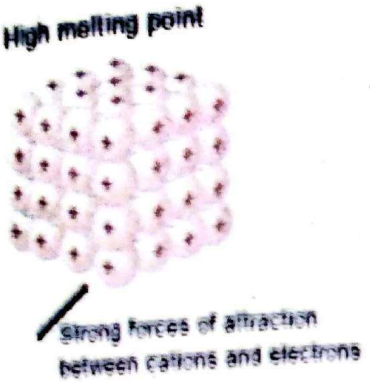
Ductility

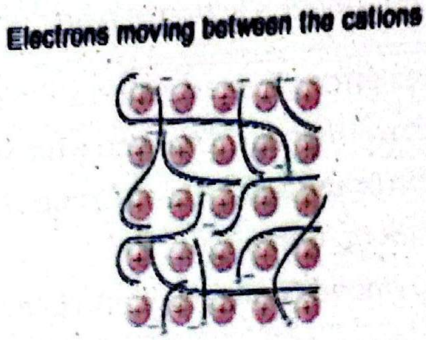
High melting and boiling point

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

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- 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	16	(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 (b) F

(c) H (d) N

(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.

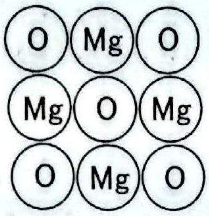
7. Explain with examples single, double and triple covalent bond.

8. Represent the formation of cations for the following metal atoms using electron dot diagram.

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

THINK TANK

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



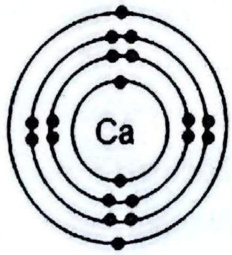
(a) What is the charge on these ions?

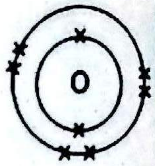
(b) How these ions get these charges?

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

11. 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.





12. 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?

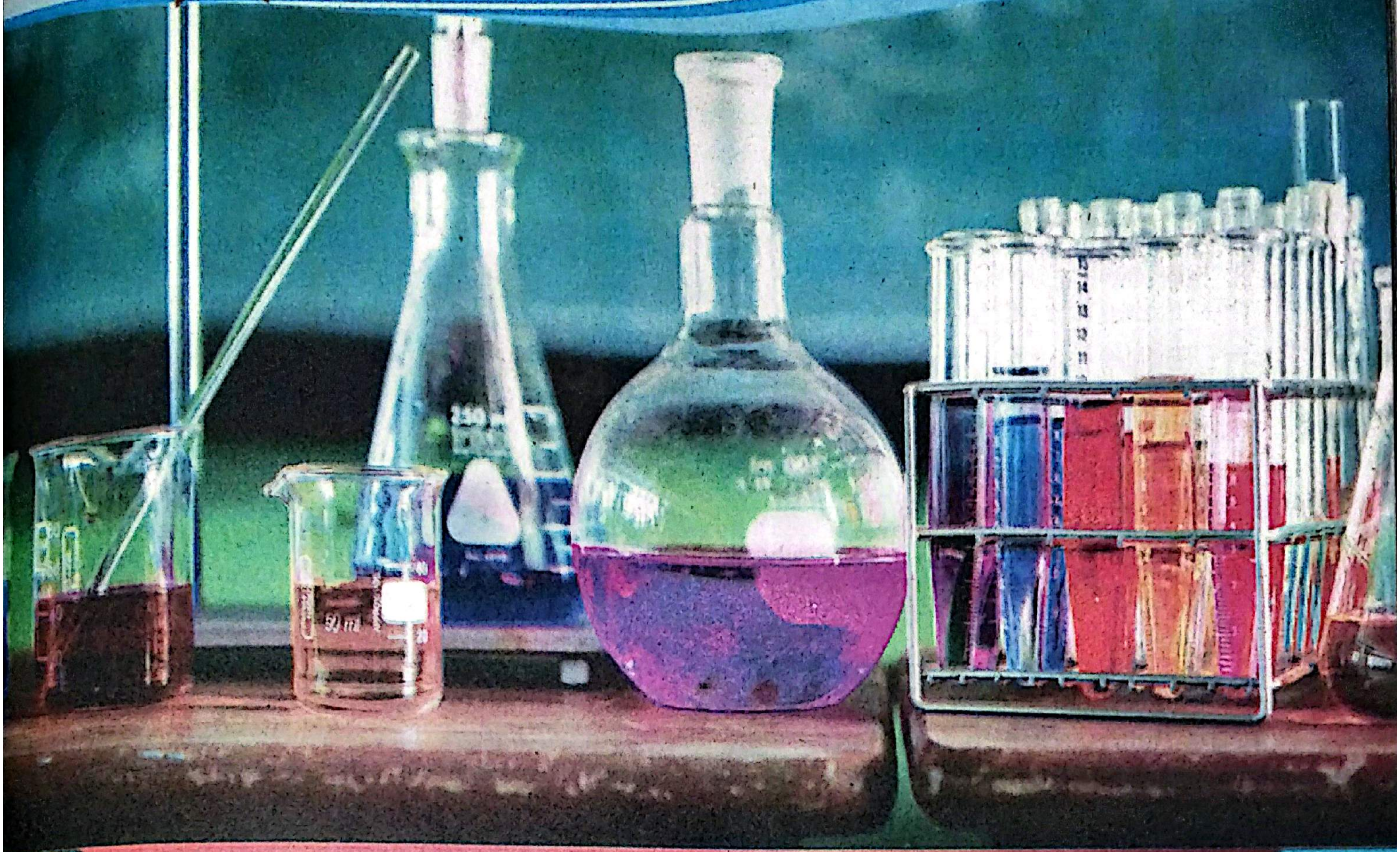
(c) Which substance is a covalent compound?

(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 , 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

= 18.016 amu

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 x 10 + 1 x 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 x Atomic mass of Na + 1 x Atomic mass of Cl
   = 1 x 23+ 1 x 35.5
   = 58.5amu
2. Formula mass of $Mg(OH)_2$ = 24 + 16x2 + 1x2
   = 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 natural.

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

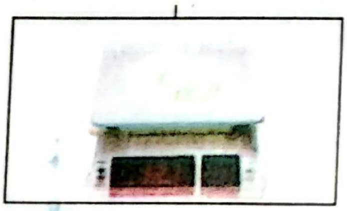
Cations		Anions	
Lithium	Li^+	Fluoride	F^-
Sodium	Na^+	Chloride	Cl^-
Ammonium	NH_4^+	Nitrite	NO_2^-
Potassium	K^+	Bromide	Br^-
Magnesium	Mg^{2+}	Nitrate	NO_3^-
Calcium	Ca^{2+}	Phosphate	PO_4^{3-}
Copper	Cu(II)	Sulfate	SO_4^{2-}

6.4 AVOGADRO'S NUMBER AND MOLE

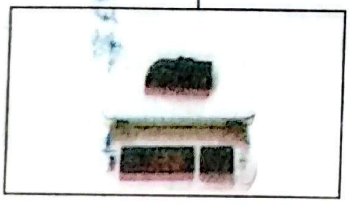
How do you count your shoes? Since the shoes come in pairs, you'll probably count them in pairs rather than individually. Likewise, eggs, oranges, etc. are counted in the dozens, but paper with the ream. So, the unit of counting depends on what you are counting. Chemists also use a practical unit for 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 represents twelve eggs, a ream of paper represents 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

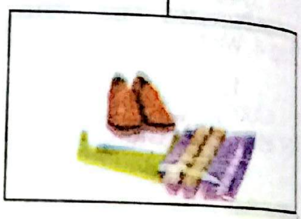


Fig. 6.1: A mole of S-atoms, a mole of C-atoms & pair of Shoes & a dozen eggs

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

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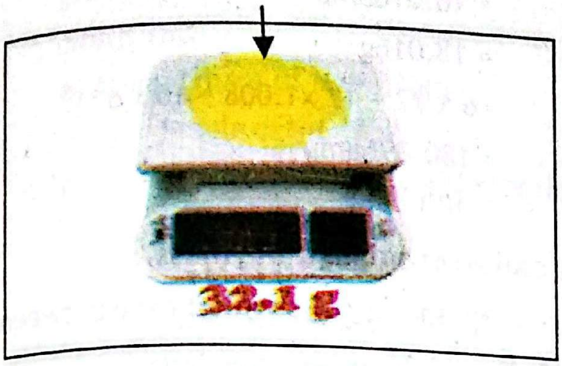
Size of the Mole

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.

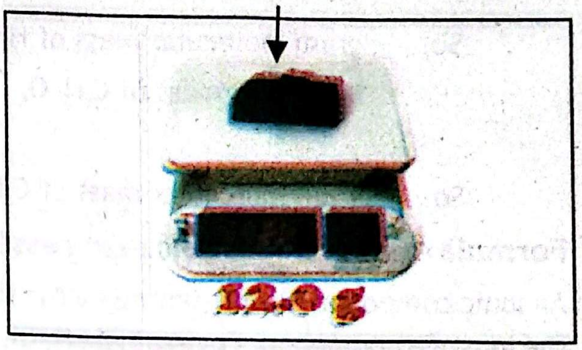
6.4.1 Gram Atomic Mass, Gram Molecular Mass and Gram Formula Mass

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.

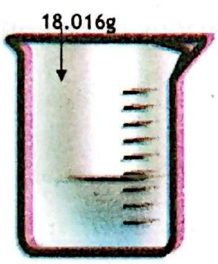
Is the gram atomic mass of C- atoms 12 g?
What is the gram atomic mass of S-atoms?
If each of the carbon and sulphur sample shown above contains one mole of atoms, why do the samples have different masses?

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

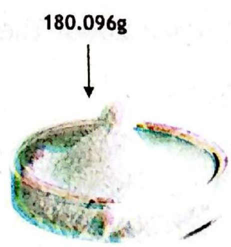
Gram atomic mass of an element contains 1 mole of atoms.
Therefore,

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

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
	= 58.5amu
Therefore, gram formula mass of NaCl	= 58.5
Formula mass of KCl	= 39 + 35.5
	= 74.5amu
So, gram formula mass of KCl	= 74.5g

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) Atomic mass of Na= 23amu

1 mole of Na = 23g

b) Nitrogen occurs as a diatomic molecules.

Molecular mass of N_2 = 14 x 2

= 28 amu

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 = 342 amu

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

$$0.25 \text{ moles of } CO_2 \longrightarrow ?\text{ g of } CO_2$$

Solution:

Molar mass of CO_2 $= 12 + 16 \times 2$

$= 44g$

1 mole of CO_2 $= 44g$ of CO_2

So, 0.25 moles of CO_2 $= 44 \times 0.25$

$= 11g$ 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 gas.

(b) A block of ice that weighs 100g.

Problem solving strategy:

Hydrogen gas and ice both are molecular substances. Determine their molar masses. Use the molar mass of each to convert the masses in grams to moles.

$$\text{mass} \longrightarrow ?\text{ moles}$$

Solution:

a) Molar mass of H_2 $= 1.008 \times 2$

$= 2.016g$

1 mole of H_2 $= 2.016g$

So, 2.016 g of H_2 $= 1$ mole of H_2

1 g of H_2 $= \frac{1}{2.016}$ moles of H_2

5 g of H_2 $= \frac{1}{2.016} \times 5$ moles of H_2

$= 2.48$ moles of H_2O

b) 1 mole of H_2O $= 2 \times 1.008 + 16$

$= 2.016 + 16$

$= 18.016g$

1 mole of H_2O $= 18.016g$

So, 18.016 g 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 x 1023 atoms = 1 mole

3.011 x 1023 atoms ⟶ ? moles

Solution:

6.022 x 1023 Ti atoms = 1 mole of Ti

$1 \text{ Ti atom} = \frac{1}{6.022 \times 10^{23}}$ moles of Ti

$3.011 \times 1023 \text{ Ti atoms} = \frac{1}{6.022 \times 10^{23}} \times 3.011 \times 10^{23}$ 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 ⟶ ? 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 x 1022 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^{30} 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 and 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 change **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 atomic number.
- A compound consists of two or more elements held together in fixed proportions 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
(c) 249.5 (d) 149.5

(ii) A compound with chemical formula Na_2CX_3 has formula mass 106 amu. 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 gas?

(a) 1 (b) 0.5
(c) 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) mass
(c) molecules (d) atoms

(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

Give short answer.

2.
(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) 250 mg of carbon
(c) 15 g of sodium chloride
(d) 40 g of sulphur
(e) 1.5 kg of MgO

6. Calculate the mass in grams of each of the following samples:

(a) 1.2 moles of K
(b) 75 moles of H_2
(c) 0.25 moles of steam
(d) 1.05 moles of $CuSO_4.5H_2O$
(e) 0.15 moles 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
(c) 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. Its molecule contains 7 C-atoms, 5 H-atoms, 3-Natoms and 6 O-atoms. Write its empirical 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 133.5 amu.

(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) 23 g of Na

(c) 5 g 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) $Na_{(s)} + H_2O_{(l)} \qquad NaOH_{(aq)} + H_{2(g)}$

(b) $NH_{3(g)} \qquad N_{2(g)} + H_{2(g)}$

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-} and oxide ion is O^{2-}.

(a) Predict the formula of potassium oxide and potassium nitride.

(b) A 0.5 g sample of K was added in 100 cm^3 of water.

$K_{(s)} + H_2O_{(l)} \qquad KOH_{(aq)} + H_{2(g)}$

Show that 1.28×10^{-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.

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 aquous 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

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 in 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

Society, Technology and Science

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- 1	+1
Group- 2	+2
Group- 13	+3
H	+1 (except in metal hydrides where it is -1)
Group- 17	-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-1. 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:

Suppose oxidation state of C = x

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

$$x - 6 = -2$$

$$x = 6 - 2$$

$$x = 4$$

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 oxidized whereas Cl_2 is oxidizing agent as it is reduced.

Activity 7.1

Prepare solutions of ferrous sulphate ($FeSO_4$) and potassium permanganate ($KMnO_4$) in separate beakers. Transfer about 10 cm^3 of ferrous sulphate solution in a test tube. Add about 10 cm^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^- 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.

I 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.

Society, Technology and Science

Redox Reaction 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.

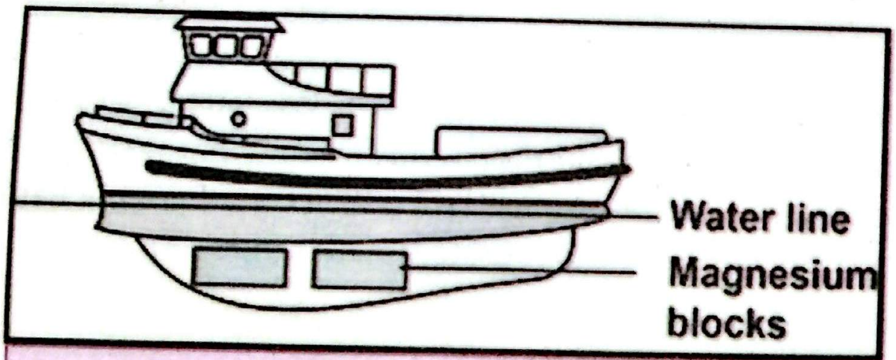


Fig. 7.1: Cathodic Protection

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 reaction changes 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

$$CuO + H_2 \longrightarrow Cu + H_2O$$

(iv) Why tin plated steel is used to make food cans?

(v) 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.

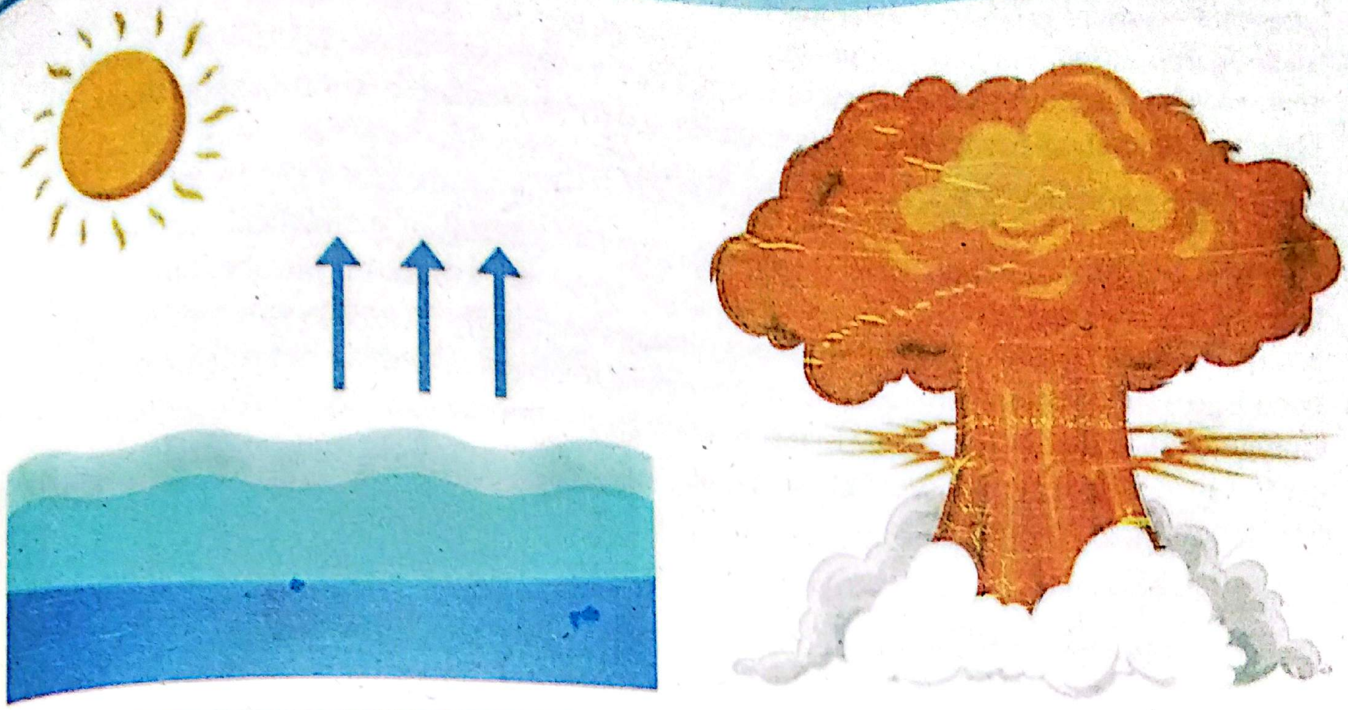
(i) $H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2$

(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

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H^{\circ} = -393.5\ kJ$$

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) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H^\circ = -393.5\ kJ$

(ii) $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$ $\Delta H^\circ = -571.6\ kJ$

(iii) $C_{(s)} + \frac{1}{2}O_{2(g)} \longrightarrow CO_{(g)}$ $\Delta H^\circ = -110.5\ kJ$

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) $H_{2(g)} + I_{2(g)} \longrightarrow 2HI_{(g)}$ $\Delta H^\circ = +53.8\ kJ$

(ii) $C_{(s)} + H_2O_{(g)} \longrightarrow CO_{(g)} + H_{2(g)}$ $\Delta H^\circ = +131.4\ kJ$

(iii) $N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)}$ $\Delta H^\circ = +180.5\ kJ$

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) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H^\circ = -393.5\ kJ$

(ii) $H_{2(g)} + I_{2(g)} \longrightarrow 2HI_{(g)}$ $\Delta H^\circ = +53.8\ kJ$

Which of the above reaction is endothermic?

8.4 BOND FORMATION 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 formation energy. The difference between the bond dissociation energy and the bond formation 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 formation 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 formation energies of products

Solution:

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}\ \Delta H^\circ = ?$

$$\Delta H^\circ = [\text{BDE of H—H} + \text{BDE. of I—I}] - [2 \times \text{BFE of H—I}]$$

$$= [436 + 151] - [2 \times 299]$$

$$= 587 - 598$$

$$= -11\ \text{kJ/mol}$$

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

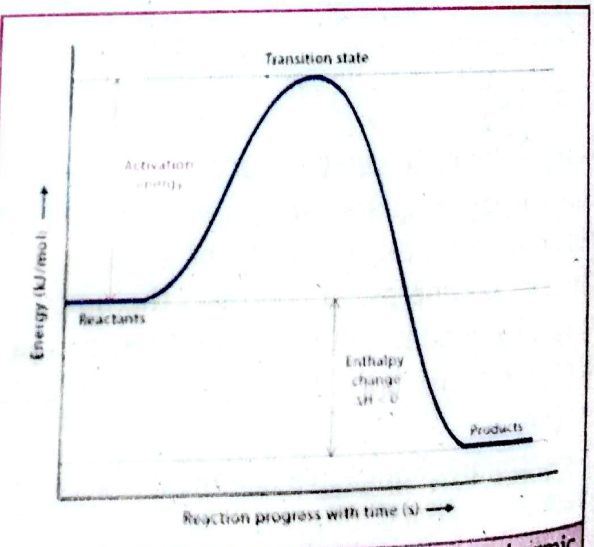


Fig. 8.1: Energy level diagram for exothermic reaction

its activation energy. The higher the activation energy, the lower the reaction rate. For example, the reaction between A_2 and B_2 molecules.

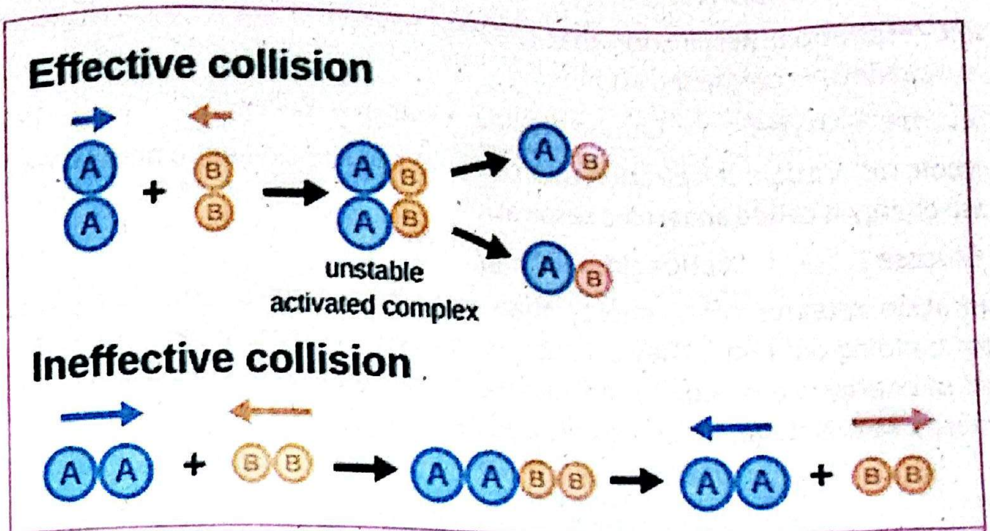


Fig. 8.2: Representation of Effective and Ineffective collision in a reaction

8.6 CATALYST

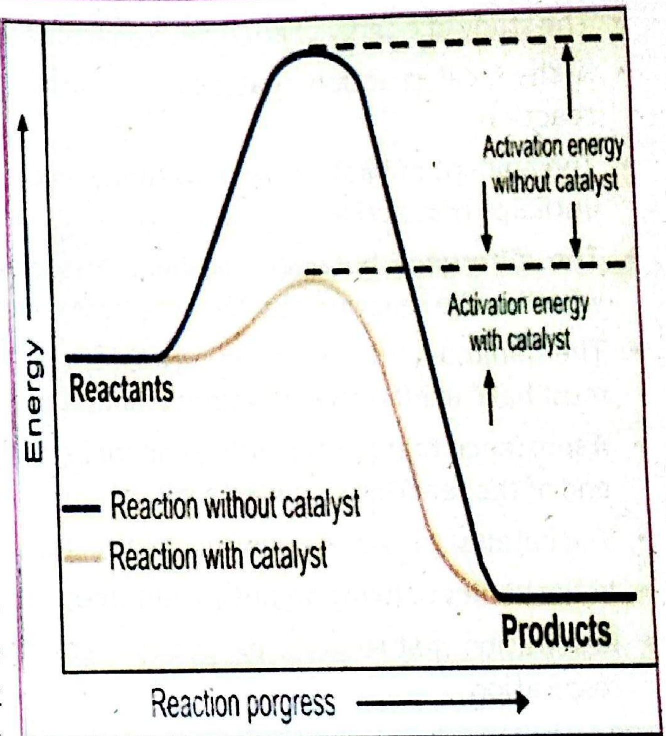


Fig. 8.3: Effect of catalyst on energy of activation during exothermic reaction

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.

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

Encircle the correct answer.

1. (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

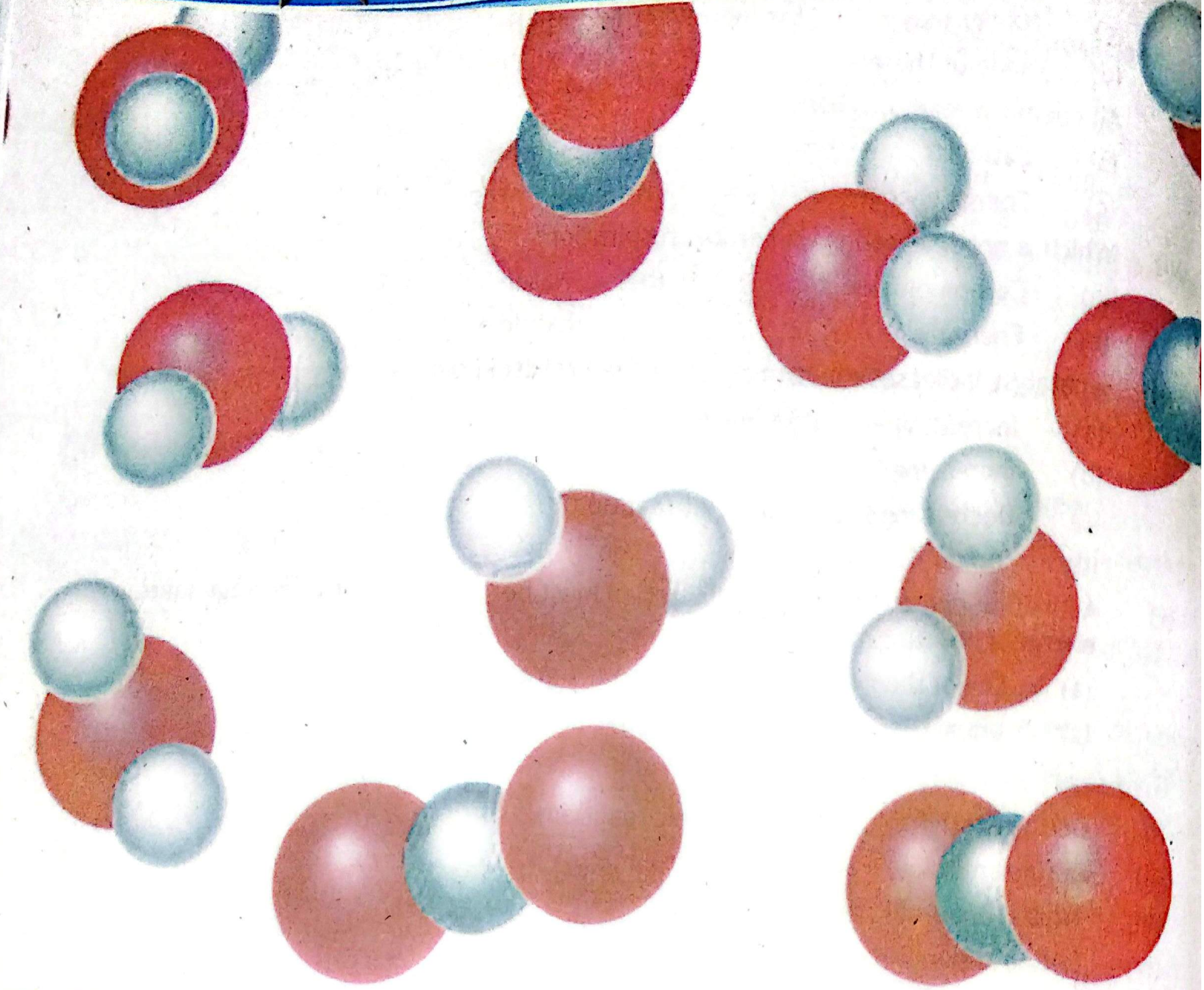
   (v) Activation energy of a chemical reaction must be __________ the average kinetic energy of reacting molecules:
   - (a) Lower than
   - (b) greater than
   - (c) equal to
   - (d) None of these

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, 243 kJ/mol and 432 kJ/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 condense as a result of collision with the surface of the liquid. Finally, the rate of evaporation equals the rate of condensation. At this stage, an equilibrium 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)

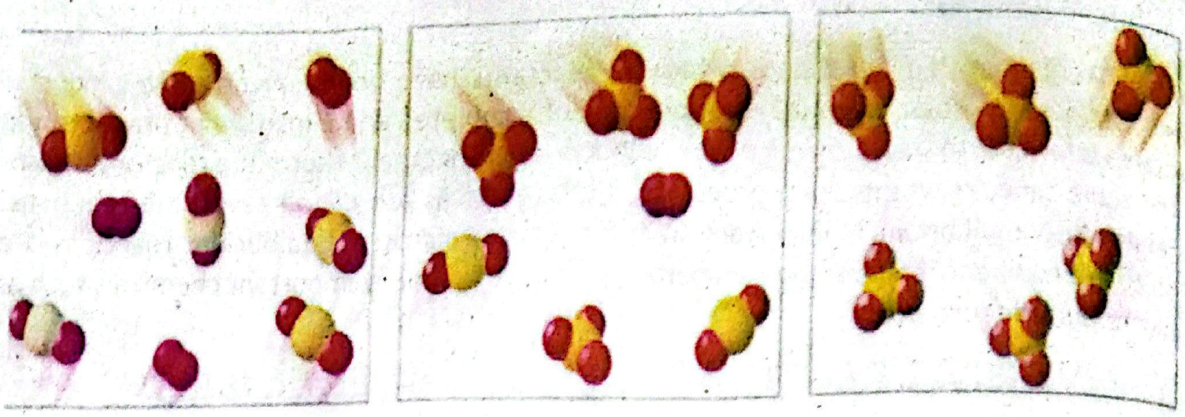


Fig. 9.1: Reaction between SO_2 and O_2

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

When the temperature is increased the equilibrium will shift to favour the reactions which will decrease the temperature. So , endothermic reaction is favoured to minimised the change.

Activity:

Materials required Test tube, dropper, heating system, hydrated copper(ll)sulphate.

Procedure:

- Place 5 g of hydrated copper(ll) 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.

$$CuSO_4.5H_2O \text{ (blue solid)} \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} CuSO_4\text{(white solid)} + 5H_2O$$

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

$$CoCl_2.6H_2O\text{(pink solid)} \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} CoCl_2\text{(blue solid)} + 6H_2O$$

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 $\rightleftharpoons$.
- 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 H_2 and I_2 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.

Give short answer.

2. (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.

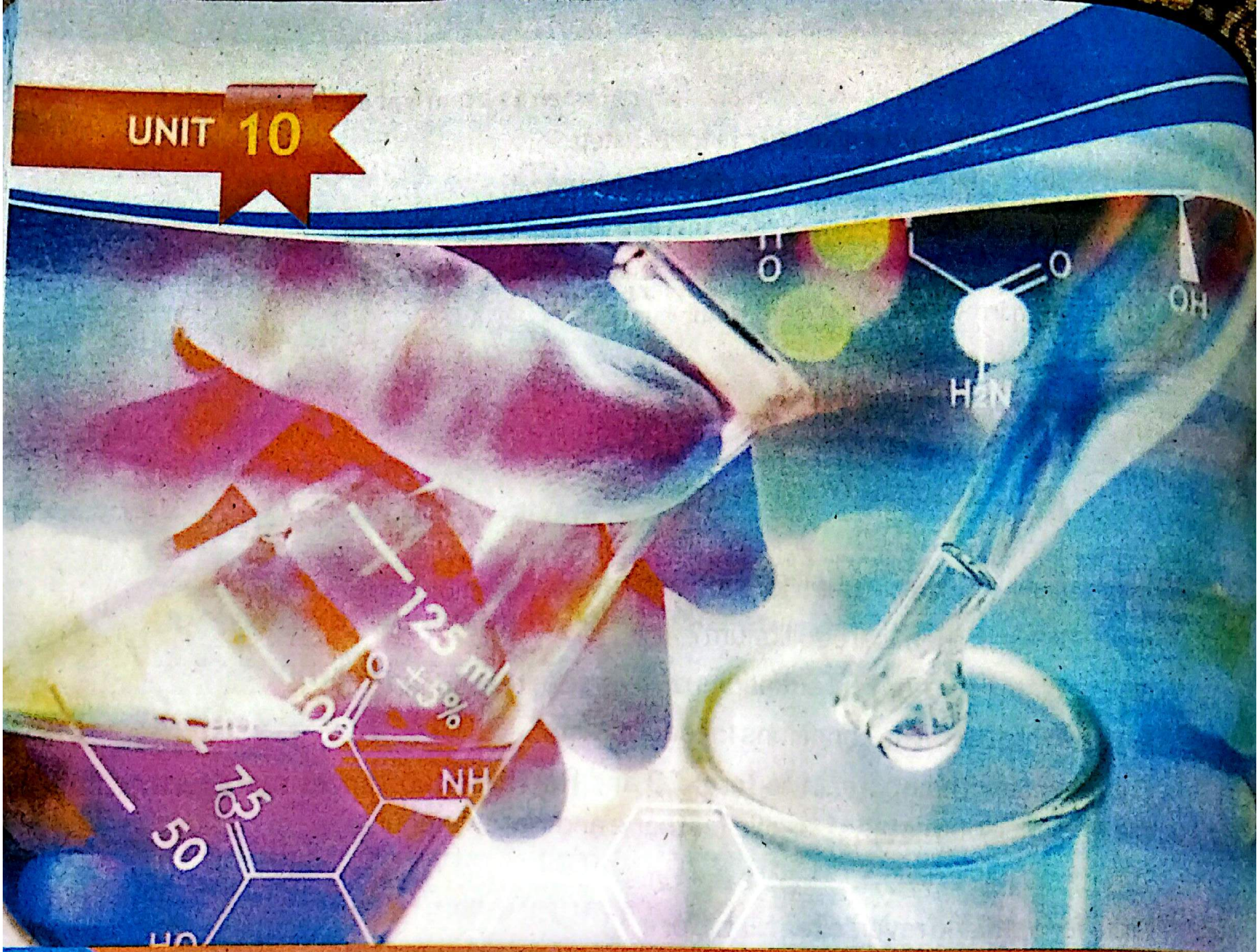
$$CO_{(g)} + 3H_{2(g)} \rightleftharpoons CH_{4(g)} + H_2O_{(g)}$$

Write forward and reverse reactions for it.

4. How does temperature affect hydrated cobalt chloride and anhydrous 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,

$$\underset{\text{Acid}}{HCl_{(g)}} \xrightarrow{H_2O} H^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$\underset{\text{Base}}{NaOH_{(s)}} \xrightarrow{H_2O} Na^+{}_{(aq)} + OH^-{}_{(aq)}$$

Which substances in the following reactions are acids or bases?

$$HNO_{3\,(l)} \xrightarrow{H_2O} H^+_{(aq)} + NO_3^{-1}{}_{(aq)}$$

$$H_2SO_{4\,(l)} \xrightarrow{H_2O} 2H^+_{(aq)} + SO_4^{-2}{}_{(aq)}$$

$$KOH_{(s)} \xrightarrow{H_2O} K^+_{(aq)} + OH^-_{(aq)}$$

$$NH_4OH_{(aq)} \xrightleftharpoons{H_2O} NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

Table 10.2 shows some common acids and table 10.3 shows some common bases.

Table 10.2: Some Common Acids names, formula and their uses

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 names, formula and their uses

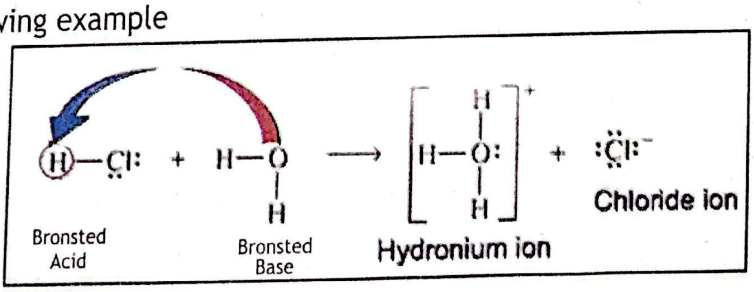
Name	Formula	Common use
Sodium hydroxide	$NaOH$	Soap making, drain cleaners
Potassium hydroxide	KOH	Making liquid soap, shaving cream
Calcium hydroxide	$Ca(OH)_2$	Making mortar, plasters, cement
Magnesium hydroxide	$Mg(OH)_2$	Antacid, laxative

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 OH 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. Which substance is donating proton?

____

Q. 2. Which substance is accepting proton?

____

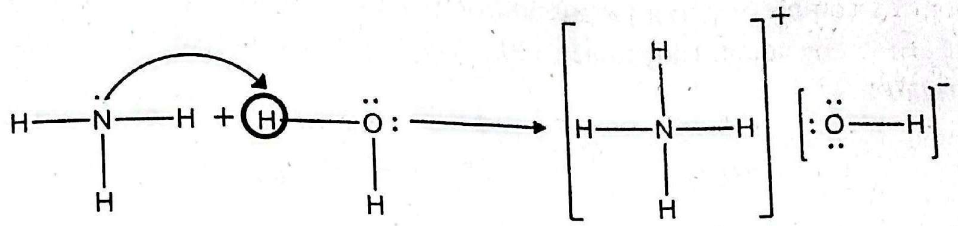
Q. 3. Which substance is an acid?

____

Q. 4. Which substance is a 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.

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$

$$H_2O + NH_3 \rightleftharpoons NH_4^+ + OH^-$$

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^+, water is 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. $H_2SO_4 + H_2O \longrightarrow HSO_4^- + H_3O^+$

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 ionizes 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, ethanoic acid(acetic acid) which is found in vinegar ionizes only up to 5% in water. So, ethanoic acid 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 a 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.

$$H_2O + NH_3 \rightleftharpoons NH_4{}^+ + OH^-$$

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($Cu(OH)_2$, aluminiun 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 taste.
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 gas. The hydrogen gas is liberated in the form of bubbles. Zinc reacts with hydrochloric acid to produce zinc chloride with the liberation of hydrogen gas. Similarly, magnesium reacts with sulphuric acid to produce magnesium sulphate and hydrogen gas. The reactions of metals with acids can be described by the following equations.

   Metal + Acid □ Salt + Hydrogen gas

   For example:

   $$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
   $$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$$

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 reactions of metal carbonate with acids can be represented by the following general reaction.

   Metal carbonate + Acid → Salt + Water + Carbon dioxide

   For example:

   $$Na_2CO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$
   $$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$$

   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 reactions can be represented by the following general reaction.

   Acid + Base → Salt + Water

   For examples:

   $$HCl + NaOH \rightarrow NaCl + H_2O$$
   $$H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$$

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.

   For Example:

   H_2SO_4 + NaOH □ Na_2SO_4 + H_2O

5. Bases decompose ammonium salts on heating and liberate ammonia gas.

   For Example:

   Base + Ammonium salt → Salt + Ammonia + 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).

Metal + Oxygen → Metal oxide

For example:

$4Na + O_2 \rightarrow 2\,Na_2O$

$2\,Mg + O_2 \rightarrow 2\,MgO$

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 the following reaction?

$O^{2-} + H\text{-}OH \rightarrow 2OH^{1-}$

Metal oxide + water → Metal hydroxide

For example:

$MgO + H_2O \rightarrow Mg(OH)_2$

$Na_2O + H_2O \rightarrow 2\,NaOH$

$CaO + H_2O \rightarrow Ca(OH)_2$

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 and 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_2... alunimium hydroxide $Al(OH)_3$, 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.

$H_2O + HCl \rightleftharpoons H_3O^+ + Cl^-$

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 ⟶ magnesium sulphate + water.

(ii) Sulphuric acid + Sodium Carbonate ⟶ Soduimsulphhate + water.

(iii) Hydrochloric acid + Magnesium ⟶ 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) $HCO_3^-{}_{(g)} + H_2O_{(l)} \rightleftharpoons CO_{3(aq)}^{-2} + H_3O^+_{(aq)}$

(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.

(a) $HI_{(aq)}$ (b) $HNO_{2(aq)}$

9. Write equations showing the ionization of the following as Arrhenius acids.

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.