بسم اللہ الرحمن الرحیم

(In the name of ALLAH the most beneficent and the most merciful)

Based on National Curriculum 2022-23

Textbook of

CHEMISTRY

GRADE 9

یہ کتاب محکمہ تعلیم حکومت بلوچستان کی جانب سے تعلیمی سال
2026 کیلئے مفت تقسیم کی جارہی ہے اور ناقابل فروخت ہے

Balochistan Textbook Board, Quetta.

Message of Chairman

Balochistan Textbook Board, Quetta is striving hard to ensure the development and publication of quality textbooks in line with the approved Curriculum for the students from Grade ECE (Primer) to XII. These textbooks are the result of intellectual contribution of renowned educationists, scholars, subject specialists and teachers of the province. Despite our constant endeavors, there are chances of mistakes, inadvertent errors and space for improvement in the textbooks. We always look forward to constructive feedback from students, teachers, parents and society at large.

The textbook in hand, has been developed by Textbook Board based on National Curriculum of Pakistan 2022-23 and is being printed as Pilot Edition for a period of one academic year i.e 2026. The feedback received will be referred to the Research Center established at Balochistan Textbook Board for analysis and drawing conclusions. The feedback after analysis will be shared with concerned Desk Officers before next edition. we are highly thankful for your valuable feedback. This research-based feedback analysis will prove to be catalyst in the improvement of quality of textbooks.

Feedback regarding the provision of textbooks at school level can also be shared.

Feedback can be shared through the following methods:

1. **website: www.btbb.com.pk**
2. **Official app on playstore "btbb official"**
3. **email btbb_quetta@yahoo.com**
4. **phone # 081-2470 501, 081-247 0503**

Dr. Gulab Khan Khilji

Chairman

Balochistan Textbook Board, Quetta.

Authors:

Mr. Abdullah Jan Zeerak,
Associate Professor of Chemistry, GPGSC Quetta.

Mr. M. Shoaib Khan Achakzai
Lecturer

Editor:

Qayyum Ullah Khan
SS (Science),

Internal Review Committee

- **Dr. Saleem Nawaz**
- **Dr. Awais**
- **Mr. Muhammad Amin**
- **Mr. Aman Ullah**
- **Qayyum Ullah Khan**

Provincial Technical Review Committee

- **Dr. Saima Jadoon**
- **Dr. Farman Ali Khan**
- **Mr. Muhammad Iqbal**
- **Dr. Afia Iqbal**
- **Dr. Syed Jamal Shah**

Printing Supervision: Mrs. Samina Mangel

Layout Designing: Farrukh Iqbal

Prepared by: New College Publications, Quetta.

Printer: Ishaq Al Fateh Printers

CONTENTS

UNIT
1

NATURE OF SCIENCE IN CHEMISTRY

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

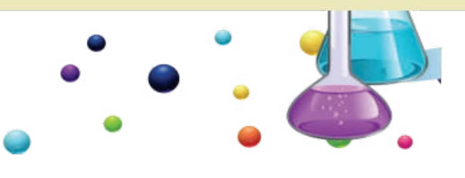
A-01 Define chemistry as the study of matter, its properties, composition, structure and interactions with other matter and energy. Or Study of earth (solids), Air (gasses), Sea (liquids) and sky (plasma) and their interaction with each other.

A-02 Explain with examples that chemistry has many sub-fields and interdisciplinary fields. (Some examples include:

- Biochemistry
- Polymer Chemistry
- Environmental Chemistry
- Organic Chemistry
- Nuclear Chemistry
- Medicinal Chemistry
- Geochemistry
- Physical Chemistry
- Inorganic Chemistry
- Astrochemistry)

A-03 Formulate examples of essential questions that are important for the branches of Chemistry. (e.g. for Analytical Chemistry a question would be 'how can we accurately determine the chemical composition of a sample?')

A-04 Differentiate between 'science', 'technology' and 'engineering' by making reference to examples from the physical sciences. (Science is a process of exploring new knowledge methodically through observation and experiments, technology refers to the process of applying scientific knowledge in practical applications for various purposes. Engineering is the application of knowledge in order to design, build and maintain a product or a process that solves a problem and fulfils a need. Science provides the foundational knowledge and understanding while engineering applies that knowledge to develop practical solutions.)



Introduction

Chemistry is the study of matter, including its composition, properties, and interactions with other matter and energy. A person (scientist) who understands chemistry is known as a chemist. Chemists work in various subfields, such as physical, biochemistry, organic, inorganic, and analytical chemistry.

Chemistry is crucial for understanding the natural world, including Earth (solids), air (gases), sea (liquids), and sky (plasmas). Chemists study the composition, structure, and changes of matter in various environments, including minerals in the Earth, atmospheric gases, properties of seawater, and ions and molecules in plasma states. While these subjects connect with geology, meteorology, and oceanography, chemistry explains the interactions and transformations of matter within these systems. This interdisciplinary approach enhances our understanding of and ability to address environmental challenges on our planet.

Do You Know?

Plasma, sometimes called the "fourth state of matter," makes up over 99% of the visible universe, including stars and lightning.

1.1 Subfields and Interdisciplinary Fields of Chemistry

The field of chemistry includes various subfields and interdisciplinary disciplines. Subfields are specialised branches within a single discipline, including organic, inorganic, and analytical chemistry. Conversely, interdisciplinary fields combine knowledge and methods from various disciplines to address complex problems. Examples of interdisciplinary fields in chemistry include biochemistry, physical chemistry, and environmental chemistry.

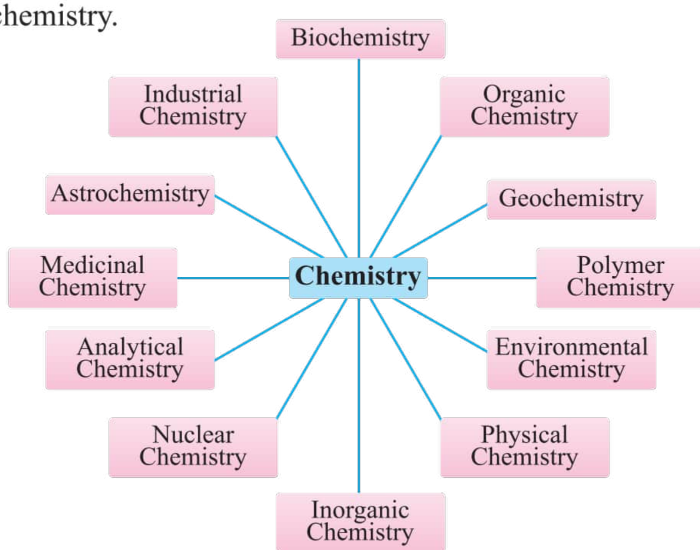


Fig 1.1: Branches of chemistry

1. **Biochemistry:** Studies the structure, composition, and chemical reactions of substances found in living organisms. It focuses on biomolecules, including carbohydrates, proteins, lipids, and nucleic acids, and explains metabolic processes such as energy production, biosynthesis, and disease-related changes.
2. **Medicinal Chemistry:** A specialised branch of chemistry that focuses on the design, synthesis, and development of pharmaceutical compounds. It combines organic chemistry, biochemistry, and pharmacology to study how chemical substances interact with biological systems for therapeutic purposes.
3. **Polymer Chemistry:** Focuses on the synthesis, structure, properties, and uses of polymers, which are large molecules made of repeating units (monomers). It includes both natural polymers (such as proteins, cellulose, and rubber) and synthetic polymers (like plastics, fibres, and resins).

Real-Life Application

Synthetic polymers, such as nylon, are used in parachutes, ropes, and medical sutures due to their exceptional strength and flexibility.

4. **Geochemistry:** The study of the chemical composition of Earth, focusing on how elements are distributed and transported within rocks, soils, water, and the atmosphere. It explains geological processes through chemical principles and the cycles of elements.
5. **Environmental Chemistry:** Examines the chemical composition of the environment and the effects of human activities on it. It includes the study of pollutants, toxic substances, and chemical processes in air, water, and soil, to protect and improve environmental quality.
6. **Analytical Chemistry:** Deals with the separation, identification, and quantification of chemical components in a sample. It involves qualitative analysis (what substances are present) and quantitative analysis (how much of each is present) using various techniques and instruments.
7. **Physical Chemistry:** Applies the laws and theories of physics to study the structure of matter and the changes it undergoes. It includes the behaviour of gases, liquids, solids, thermodynamics, kinetics, quantum chemistry, and spectroscopy.
8. **Organic Chemistry:** Focuses on studies covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives. It includes natural and synthetic compounds, with applications in petroleum, petrochemicals, and pharmaceuticals.
9. **Inorganic Chemistry:** Deals with all elements and their compounds, except most carbon–hydrogen compounds. It has wide applications in materials such as ceramics, glass, cement, and metallurgy.

10. **Nuclear Chemistry:** Explores radioactivity, nuclear reactions, and the chemical effects of radiation. It includes atomic energy, nuclear reactors, radiotherapy, food preservation, and isotopic studies.
11. **Astrochemistry:** The study of the chemical composition and reactions of matter in outer space, including stars, planets, interstellar clouds, and cosmic dust. It connects chemistry and astronomy to explain the origin of elements and molecules in the universe.

Chemistry in Your Body

Calcium ions (Ca^{2+}) not only make up part of your bones and teeth, but also play an essential role in muscle contraction and nerve signalling.

Essential Questions

Biochemistry: Why is water essential for living organisms?
Medicinal Chemistry: How do medicines support our bodies? Name one common medicine you know.
Polymer Chemistry: Provide two common examples of polymers.
Geochemistry: What is the Earth made of? Why is it important to study the Earth's chemicals?
Environmental Chemistry: What is the greenhouse effect? Name two greenhouse gases.
Analytical Chemistry: What standard techniques are used to determine the chemical composition of a sample?
Physical Chemistry: How does temperature affect the rate of a chemical reaction?
Organic Chemistry: What is a hydrocarbon? Provide an example.
Inorganic Chemistry: Identify essential inorganic elements found in the human body.
Nuclear Chemistry: How does a nuclear reaction differ from a chemical reaction?
Astrochemistry: What is the composition of space? Can chemicals exist in space?
Industrial Chemistry: How is soap produced? Write a simple word equation for the saponification process.

COMMON MISTAKE

Students often confuse physical and chemical changes. Melting of ice is a physical change, but rusting of iron is a chemical change.

1.2 Science, Technology, and Engineering

Science, technology, and engineering are interconnected, working together to help us understand and improve the world around us.

1.2.1 Science

Science is the systematic process of discovering new knowledge through observation, experimentation, and analysis. It aims to understand and explain the principles and laws that govern the natural world, living organisms, and natural phenomena. Scientists formulate scientific theories by conducting experiments and collecting and analysing data to develop hypotheses and draw conclusions. Various scientific fields include chemistry, physics, astronomy, and environmental science.

Applications of Science

In science, physics helps us understand concepts like electromagnetic waves, which enable us to send information through the internet, phone calls, and videos. Newton's Laws of Motion guide spacecraft and missiles, ensuring they travel accurately. Atomic energy is used to produce clean and efficient power. Chemistry creates new materials, such as strong composites and plastics, and is essential for developing medicines that treat diseases. Astronomy allows us to study stars and planets with high-resolution tools and employs satellites for communication and navigation.

1.2.2 Technology

Technology involves applying scientific knowledge and tools for practical purposes, aiming to solve problems and enhance our lives. It includes the creation of physical objects, new materials, systems, and devices that serve specific functions. This may use existing tools and techniques in innovative ways. Technology can be applied across various fields, including engineering, social sciences, medicine, agriculture, and architecture.

Applications of Technology

Technology transforms scientific ideas into practical tools. For example, nanotechnology employs tiny materials to create innovative devices and medicines. Renewable energy technologies use solar panels, wind turbines, and hydroelectric plants to generate clean energy. Medical imaging technologies, such as X-rays and MRI scans, enable doctors to visualise the inside of the body for diagnosing illnesses.

1.2.3 Engineering

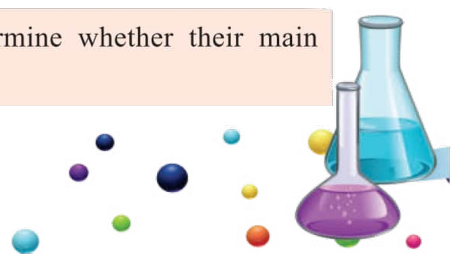
Engineering involves the design, construction, testing, and maintenance of structures, machines, and systems to solve problems and fulfil specific requirements. Engineers apply principles of science, mathematics, and creativity to create, develop, and implement solutions for a wide array of real-world challenges. Numerous engineering disciplines, including civil, software, mechanical, and electrical engineering, play an essential role in various industries and applications.

Applications of Engineering

Engineering applies science and technology to build and enhance machines and systems. Mechanical engineering focuses on designing engines and machines that transform energy into work. Electrical engineering specialises in developing electrical systems for power and communication. Materials engineering creates new materials with unique properties for use across various industries, from construction to aerospace.

ACITIVITY

Collect labels from five household cleaning products. Determine whether their main components are acids, bases, or salts.



KEY POINTS

1. Chemistry is the study of matter, including its composition, structure, properties, and interactions with energy and other matter.
2. Chemists study matter in all forms: solids (earth), liquids (sea), gases (air), and plasma (sky), helping us to understand natural systems and their changes.
3. Chemistry plays a vital role in daily life and major fields such as medicine, agriculture, materials science, and environmental protection.
4. Chemistry helps understand food, health, cleaning products, fuels, plastics, and natural events like rusting or photosynthesis.
5. Science involves discovering new facts through experiments, such as testing how substances react.
6. Technology refers to the application of science to create useful tools, such as microscopes or mobile phones.
7. Engineering involves applying science and technology to design and create functional structures, such as bridges, machines, or water cleaning systems.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) Chemistry primarily focuses on the study of:
   (a) living organisms (b) matter and its properties
   (c) celestial bodies (d) electrical engineering

ii) The option that is **NOT** a subfield of chemistry is:
   (a) organic chemistry (b) environmental Chemistry
   (c) civil engineering (d) analytical chemistry

iii) The interdisciplinary field that integrates chemistry with living organisms is:
   (a) polymer chemistry (b) geochemistry
   (c) biochemistry (d) medicinal chemistry

iv) The field of chemistry concerned with studying geological materials and celestial bodies is:
   (a) geochemistry (b) medicinal chemistry
   (c) analytical chemistry (d) polymer chemistry

v) The role that environmental chemistry plays in solving problems is:
   (a) drug discovery (b) pollution control and climate change
   (c) material science (d) nuclear reactions

vi) The aspect that physical chemistry studies on the atomic and molecular level is:
   (a) earth and environmental sciences   (b) behaviour of matter
   (c) organic compounds   (d) industrial chemicals

vii) Inorganic chemistry is concerned with the study of:
   (a) living organisms   (b) organic compounds
   (c) inorganic compounds   (d) radioactive materials

viii) The study of radioactivity and nuclear reactions is known as:
   (a) biochemistry   (b) environmental chemistry
   (c) nuclear chemistry   (d) polymer chemistry

ix) The subfield of chemistry that studies pollutants and their impact on air, water, and soil is:
   (a) inorganic   (b) environmental
   (c) physical   (d) industrial

x) Geochemistry mainly studies the chemical composition of:
   (a) living tissues   (b) interstellar clouds
   (c) geological materials like rocks and soil
   (d) plastics and synthetic polymers

B. Restricted Response Questions (RRQs)

i) What does polymer chemistry study?

ii) Which branch of chemistry studies planets and stars?

iii) Which branch of chemistry helps us understand pollution?

iv) What question does analytical chemistry help answer?

v) Write short notes on the following branches of chemistry:
   a) Biochemistry   b) Environmental Chemistry
   c) Nuclear Chemistry

C. Extended Response Questions (ERQs)

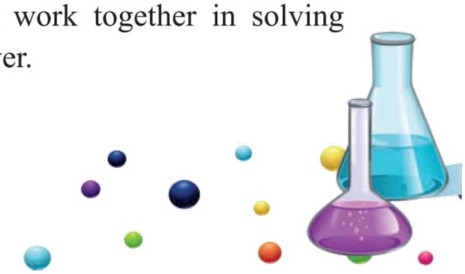
i) Explain how chemistry helps us understand the earth (solids), air (gases), sea (liquids), and sky (plasmas) with suitable examples.

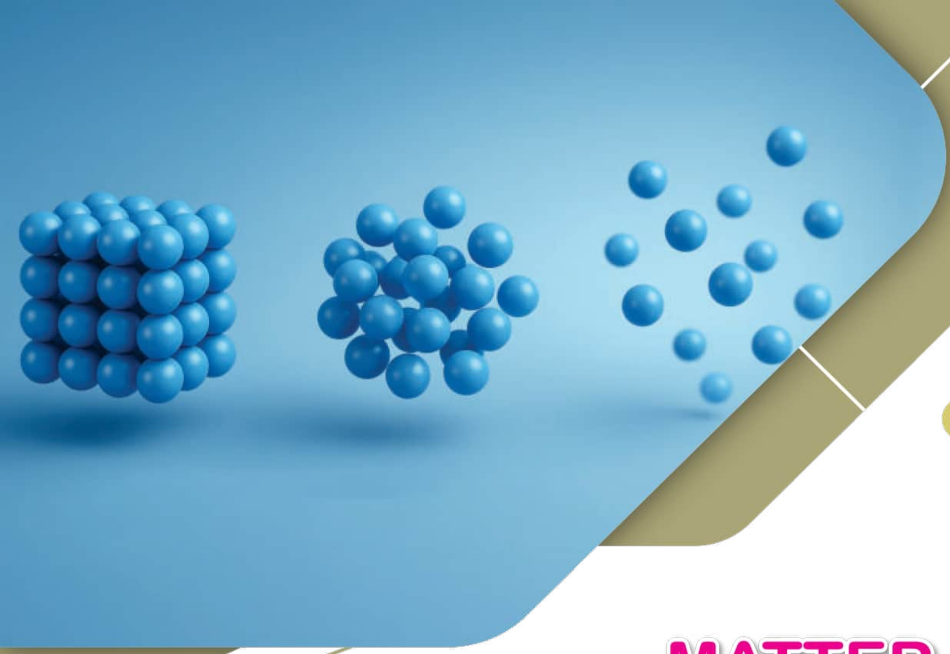
ii) Describe the difference between subfields and interdisciplinary fields of chemistry. Give two examples of each.

iii) Choose any four branches of chemistry and write one essential question that scientists working in each field try to answer.

iv) Differentiate between science, technology, and engineering. Give one clear example of each from chemistry or physical sciences.

v) Explain how science, technology, and engineering work together in solving problems. Use a real-life example to support your answer.





UNIT 2

MATTER

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-01 Define matter as a substance having mass and occupying space.

B-02 State the distinguishing macroscopic properties of commonly observed states of solids, liquids and gasses in particular density, compressibility, and fluidity.

B-03 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.)

B-04 Explain the allotropic forms of solids (some examples may include diamond, graphite, and fullerenes.)

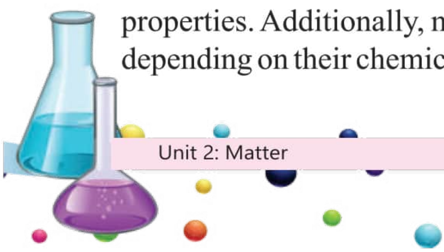
B-05 Explain the differences between elements, compounds and mixtures.

B-06 Identify solutions, colloids, and suspensions as mixtures and give an example of each.

B-07 Explain the effect of temperature on solubility and formation of unsaturated and saturated solutions.

Introduction

Matter is anything that occupies space and has mass. Water, air, trees, stones, animals, plants, carbon, silver, gold, diesel, petrol, and stars are all different types of matter because they have both mass and volume. Ancient philosophers suggested that all matter, whether living or non-living, was made up of five basic elements: earth, air, fire, water, and sky. Modern scientists classify matter based on its physical and chemical characteristics. Matter is classified into four states: solids, liquids, gases, and plasma, based on its physical properties. Additionally, matter can be classified into elements, compounds, and mixtures depending on their chemical properties.



Do You Know?

In science, emotions and ideas such as love, hate, affection, or thoughts are not classified as matter because they do not have measurable mass or volume. These are considered abstract concepts.

2.1 Macroscopic Properties of Solids, Liquids, and Gases

The study of macroscopic properties is essential for understanding the behaviour of matter in different states. These properties help us in distinguishing among solids, liquids, and gases.

2.2.1 Properties of Solids

Solid is the physical state of matter with a definite shape and volume. They may consist of pure elements and compounds. Examples include diamond, urea, ice, sucrose, paper, vegetables, fruits, wood, and rocks.

Fig. 2.1: Particles are closely packed together in a solid

Some common properties of solids include:

I) **Shapes:** Solids have definite shapes. Crystalline solids have definite geometric shapes.

ii) **Density:** The particles (atoms, ions, or molecules) of solids are very close to each other, tightly packed, more ordered than liquids, and much more ordered than gases. For that reason, solids are slightly denser than liquids and much denser than gases.

iii) **Compressibility:** Solids are generally less compressible than liquids and are often considered incompressible. While materials like foam, wood, and cork may seem compressible, they are not. These substances contain spaces filled with air. Although applying pressure compresses the spaces, the solid matter in foam, wood, or cork remains unchanged.

iv) **Fluidity:** The solid particles are locked in place, preventing their molecules from slipping and sliding over one another. Therefore, they are unable to flow (i.e., they are rigid) and thus retain their shapes and volume without a vessel.

v) **Kinetic Energy:** Solid particles have lower kinetic energy than those in liquids and gases. Kinetic energy is directly proportional to the temperature of solid materials.

vi) **Diffusion:** Solid particles do not readily diffuse into other solids. Although diffusion in solids does occur, it is very slow under high pressure.

2.2.2 Properties of Liquids

Liquids have a definite volume but lack a definite shape. Examples of liquids include water, blood, mercury, diesel, vegetable oil, and milk. Liquids are made up of atoms or molecules.

Some common properties of liquids are:

i) **Density:** Liquids have closely packed molecules, resulting in negligible empty space. Consequently, they are generally denser than gases, yet usually less dense than solids.

ii) **Compressibility:** Liquids are generally considered incompressible. They show slightly greater compressibility than solids and are more difficult to compress than gases. The space between the molecules of a liquid is greater than that of a solid and less than that of a gas.

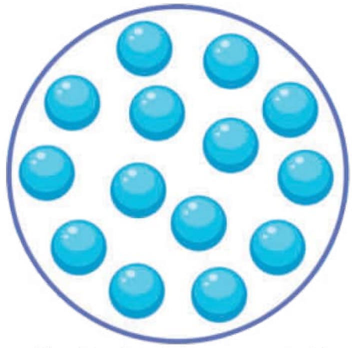


Fig. 2.2: Particles are not held in fixed positions in a liquid

iii) **Fluidity:** The liquid molecules are in a constant state of motion, that is why they slip and slide over one another. Therefore, liquids have no fixed shape and adopt the shape of the container in which they are placed.

iv) **Kinetic Energy:** Liquid molecules have higher kinetic energy than solids but lower kinetic energy than gases. Their kinetic energy decreases with a drop in temperature and increases with a rise in temperature.

v) **Diffusion:** Liquids diffuse into other miscible liquids because their molecules move around. Diffusion in liquids occurs more slowly than in gases.

QUICK CHALLENGE

If you pour 100 mL of water into a conical flask, what shape will it assume and why?

2.2.3 Properties of Gases

A physical state of matter that neither has a definite volume nor shape is known as a gas. Examples include oxygen, carbon dioxide, and methane. All gases consist of a large number of tiny particles called atoms or molecules.

Typical properties of solids include:

i) **Density:** Gases have a much lower density than solids and liquids.

ii) **Compressibility:** Gas molecules are widely separated from each other, which facilitates their easy compression.

iii) **Fluidity:** Gas molecules are in constant, random motion, colliding with each other and with the walls of their container. This property allows gases to move and adapt the shape of their containers. The fluidity of gases enables them to show properties such as diffusion, effusion, expansion, and compression.

iv) **Kinetic Energy:** Gas molecules have higher kinetic energy than liquids and solids. This is due to their high mobility, negligible intermolecular forces, and large spaces between the molecules. The average kinetic

FUN FACT

One litre of water can produce over 1,600 litres of steam at 1 atm when boiled, illustrating the significant difference in the spacing of gas molecules compared to liquids.

energy of gas molecules is directly proportional to the temperature of the gas. As the temperature increases, the average kinetic energy of the gas molecules also increases.

v) **Diffusion:** Gases can diffuse faster than liquids and solids due to their high kinetic energy, weak intermolecular forces, large spaces between molecules, constant motion, frequent collisions, and low density. Lighter gases diffuse more rapidly than heavier ones. For example, helium gas (molar mass = 2g/mol) diffuses faster than oxygen gas (molar mass = 32g/mol).

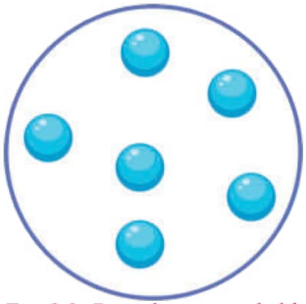


Fig. 2.3: Particles are not held in fixed positions in a gas. They move freely in all directions

Table 2.1: Comparison of properties of solids, liquids, and gases

Properties	Solids	Liquids	Gases
Shape	Definite	Indefinite	Indefinite
Volume	Definite	Definite	Indefinite
Attractive Forces	Strong	Moderate	Negligible
Particles	Tightly packed	Loosely packed	Independent
Density	High	Low	Very low
Compressibility	Incompressible	Slight	High
Fluidity	Cannot flow	Can flow	Can flow
Rigidity	High	Less	Negligible
Kinetic Energy	Low	Intermediate	High
Diffusion	Negligible	Less	High

ACITIVITY

Describe an activity to show that gases can diffuse. What does this activity prove about the nature of particles of matter?

2.3 Exotic and Intermediate States of Matter

The three most common states of matter are solid, liquid, and gas; however, several other exotic (unusual) and intermediate states exist. Each state of matter has unique characteristics determined by the arrangement and behaviour of its particles. Here are some examples representing each state:

2.3.1 Plasma

Plasma is a unique form of matter often regarded as the fourth state, distinct from solid, liquid, and gas. It is formed by heating and ionising a gas to extremely high temperatures, yielding a mixture of positively charged ions and free electrons. Unlike ordinary gases, plasma lacks a defined shape or volume and shows unique behaviours. These charged particles move freely and are influenced by electric and magnetic fields. Examples of naturally occurring plasmas include stars (including the Sun), lightning, neon lights, auroras, and flames. Scientists and researchers produce plasmas in laboratories and industries for specific applications. Common examples of artificial plasma include plasma TVs, fluorescent lights, plasma lamps, and experimental fusion reactors.

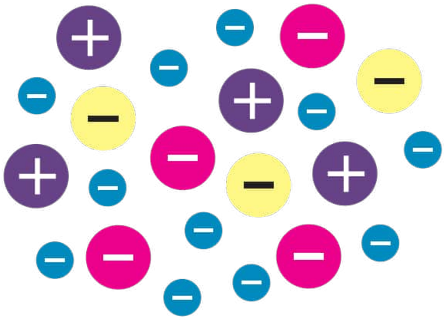


Fig. 2.4: Plasma state

2.3.2 Bose-Einstein Condensate (BEC)

Bose-Einstein condensate is a unique fifth state of matter that typically forms at extremely low temperatures, close to absolute zero (zero kelvin). In this state, a group of atoms from certain elements ceases to behave as separate particles and begins to function as a single entity. Particle movement nearly stops, and atoms start to clump together. Satyendra Nath Bose and Albert Einstein first predicted this phenomenon in the 1920s, and experimental evidence confirmed it in 1995. This represented a significant advancement in scientific knowledge in the fields of chemistry and physics, with applications in modern technologies, such as the production of super-fast computer chips. Bose-Einstein condensates (BECs) are frequently created using extremely cold gases of alkali metals (Li, Na, K, and Rb).

2.3.3 Liquid Crystals

Liquid crystals (LCs) represent a unique state of matter with properties that exist between those of a typical liquid and a solid crystal.

$$\text{Solid} \rightleftharpoons \text{Liquid Crystals} \rightleftharpoons \text{Liquid}$$

The crystalline solid has specific melting points. The temperature of solids remains constant at the melting point until all the solid has melted. However, some solids first change into a turbid liquid before transforming into a clear liquid. These turbid liquids possess properties of both solids, such as optical activity, and of liquids, such as fluidity and viscosity, and are called liquid crystals. They are frequently used in display technologies, such as LCDs (liquid crystal displays). LCDs are utilised in various electronic devices, including digital wristwatches, clocks, flat-panel televisions, computers, laptop screens, cellphones, and pH meters. Devices that utilise liquid crystals are known as LCD (Liquid Crystal Display) devices. Liquid crystals detect blockages in veins and arteries, as well as infected areas, such as tumours. Additionally, they are used in thermometers to measure the body temperature of infants.

Fig. 2.5: The liquid and liquid crystalline phases of cholesteryl benzoate

2.4 Allotropic Forms of Solids

The various forms of the same element are known as allotropes (or allotropic forms). They are composed of atoms of the same element, but their atomic arrangement varies in each structure. Despite having identical chemical properties, their physical properties differ, including hardness, density, and colour. There are three main types of carbon allotropes: diamond, graphite, and fullerenes. Other elements that exist in allotropic forms include tin, sulphur, and phosphorus. The study of allotropes is essential for understanding the properties and behaviour of various substances.

Do You Know?

In amorphous (shapeless) solids, the particles are arranged irregularly, so they have no fixed structure. They look hard, like true solids, but soften gradually rather than melting sharply. Examples include glass, wax, butter, plastics, rubber, and some types of candy.

2.4.1 Diamond

Diamond is the purest form of carbon and occurs naturally in the earth's crust. Although it is possible to create diamonds artificially, the high cost and low quality make this method uncommon. Diamonds are known to be the hardest substance, possessing high density and melting points. They are transparent and colourless, valued in jewellery for their brilliance and sparkle, and used in cutting and grinding tools due to their hardness.

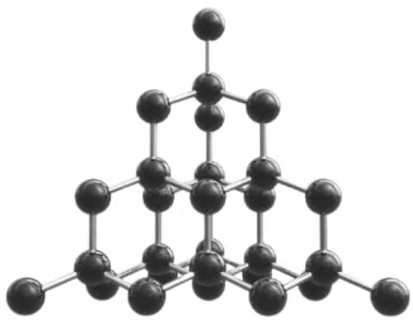


Fig. 2.6: Diamond crystal structure

Diamonds consist of a three-dimensional network of carbon atoms arranged tetrahedrally, composed of several tetrahedral units, with each carbon atom bonded to four others through strong covalent bonds. Because all four valence electrons are utilised for bonding, diamonds lack free electrons and consequently do not conduct electricity. The unique properties of diamonds, such as their ability to refract light and disperse colours, arise primarily from their crystal structure.

2.4.2 Graphite

Graphite is another form of pure carbon, having properties quite different from those of diamonds. It has a giant covalent structure composed of layers or sheets of carbon atoms stacked above one another. Each carbon atom in a layer is bonded to three other carbon atoms through covalent bonds, forming flat hexagonal rings. The various layers are held together by weak forces.

Graphite is an excellent conductor of electricity because of its layered structure. It is commonly used in lead pencils and locks due to the weak forces between its layers. The high melting point of graphite is a result of the strong covalent bonds between the carbon atoms in each layer. Furthermore, graphite appears opaque and black due to the arrangement of carbon atoms and their interaction with light.

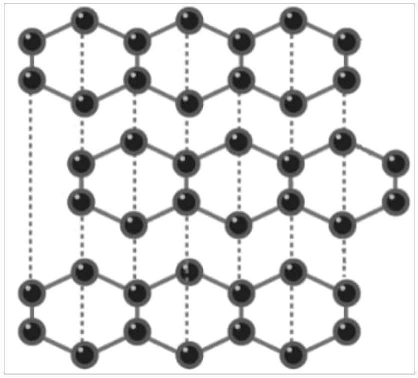


Fig. 2.7: Structure of graphite

2.4.3 Fullerene

Fullerenes (also known as Buckyballs) are allotropes of carbon that consist of molecules made entirely of carbon atoms. The most common form is C_{60}, arranged as hollow spheres, ellipsoids, or tubes. The structure of the Buckyball (C_{60}) resembles that of a soccer ball, containing 32 faces, 12 pentagons, and 20 hexagons. Each carbon atom in the molecule is covalently bonded to three other carbon atoms. Fullerene molecules are held together by weak forces. Fullerenes are rare, with small amounts of these structures (Buckyballs) produced in the forms of C_{60}, C_{70}, C_{76}, C_{82}, and C_{84} molecules. They have been discovered in soot and the residue of carbon arc lamps, and are also generated by lightning discharges in the atmosphere. Fullerenes typically do not conduct electricity; their conductivity is generally more similar to that of semiconductors or insulators, depending on the specific type and structure of the fullerene. They are commonly used in energy storage devices (such as batteries), lubricants, and coatings, and help remove contaminants from water.

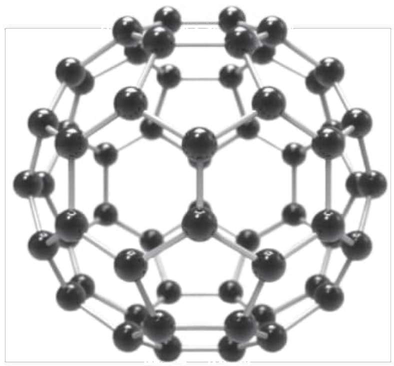
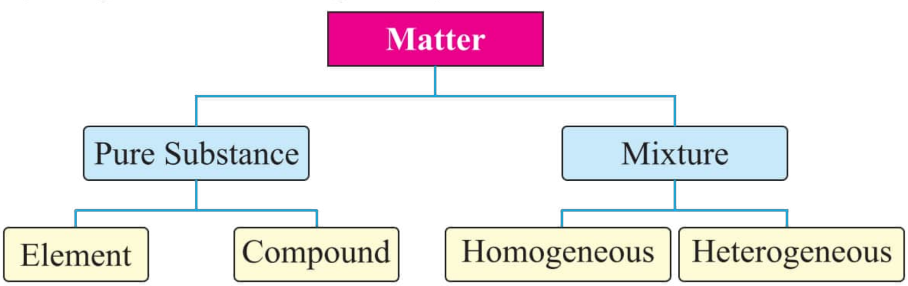


Fig. 2.8: Buckyball C_{60}

2.5 Elements, Compounds, and Mixtures

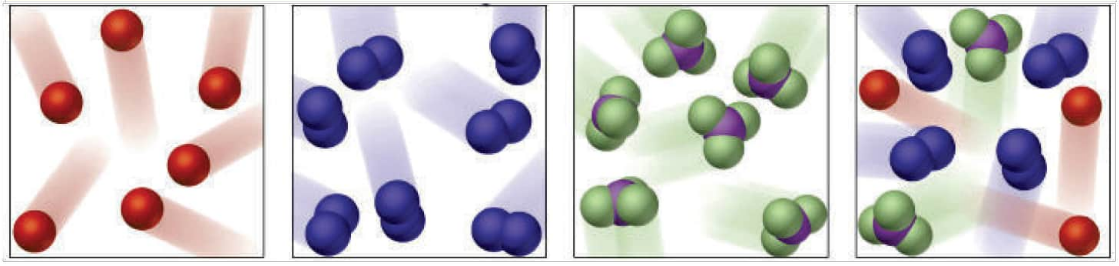
Matter is generally classified into pure substances (elements and compounds) and mixtures (homogeneous and heterogeneous).



Pure substances are types of matter composed of only one kind of particle, either atoms or molecules, and show uniform properties throughout their mass. Examples of pure substances include carbon, oxygen, silver, gold, water, and carbon dioxide. Pure substances are further classified into elements and compounds.

Table 2.2: Differences between elements, compounds and mixtures

Property	Element	Compound	Mixture
Definition	A pure substance consists of only one kind of atom.	A pure substance consists of two or more different elements in fixed ratios that are chemically combined.	An impure substance consists of two or more substances that have no chemical bonds.
Substance Category	Pure substance	Pure substance	Impure substance
Types	Metals, non-metals, and metalloids	Covalent, ionic	Homogeneous, heterogeneous
Composition	One kind of atoms	Different types of atoms (elements)	Different substances
Properties	Unique properties are determined by the type of atom	Different from the elements they contain	Same as the individual components
Separation	Cannot be broken down into simpler substances	Elements can be separated by chemical reactions or electricity	Components can be separated by physical methods
Bonding	Can form chemical bonds with other elements	Can form chemical bonds with other elements	Cannot form chemical bonds between components
Examples	Carbon, nitrogen, zinc, gold	Water, ammonia, sodium chloride	Concrete, sugar water, air



(a) (b) (c) (d)

Fig. 2.9: (a) Atoms of an element (b) Molecules of an element (c) Molecules of a compound (d) Mixture of elements and compound

2.6 Solutions, Colloids, and Suspensions

Mixtures are classified into solutions, colloids, and suspensions.

2.6.1 Solution

A solution is a homogeneous mixture of two or more components, with particles uniformly distributed throughout. It is transparent and does not scatter light. The solute particles in solutions are so small that they cannot be observed by the naked eye or a microscope.

2.6.2 Colloid

A colloid (or colloidal dispersion) is a heterogeneous mixture of two or more components. In the case of colloids, we use the terms "dispersed phase" for solute and "dispersing medium" for solvent.

2.6.3 Suspension

Suspension is also a heterogeneous mixture; the particles are large enough to be seen with a microscope and can often be observed with the naked eye. Suspensions scatter light.

(a) (b) (c)

Fig. 2.10: (a) Light passing through solution, (b) Light scattering by colloid, (c) Light scattering by suspension

Table 2.3: Properties of solutions, colloids and suspensions

Property	Solution	Colloid	Suspension
Type of mixture	Homogeneous	Heterogeneous	Heterogeneous
Size of particles	0.01nm to 1nm	1nm to 1000nm	Greater than 1000nm
Effect of light	Cannot scatter light (no Tyndall effect)	Scatter light (Tyndall effect visible)	Scatter light strongly
Transparency	Transparent	Cloudy	Opaque
Settling properties of particles	Do not settle down on standing	Do not settle down on standing	Settle down on standing
Filterability	Non-filterable through ordinary filter paper	Non-filterable through ordinary filter paper	Filterable through ordinary filter paper
Visibility of particles	Invisible to the naked eye and the microscope	Invisible to the naked eye and under an ordinary microscope	Visible to the naked eye
Examples	Salt solution, sugar solution, vinegar in water, petrol, air, soft drinks	Milk, starch solution, fog, smoke, clouds, toothpaste,	Muddy water, sand in water, flour in water, soot, milk of magnesia

Self-Assessment

Classify the following as a solution, colloid, or suspension:

i)	Milk	ii)	Vinegar in water
iii)	Muddy water	iv)	Fog

2.7 Effect of Temperature on Solubility

Solubility refers to the ability of a substance (known as the solute) to dissolve in a specified amount of another substance (referred to as the solvent) under specific temperature and pressure conditions. It is typically expressed in grams of solute per 100 grams or per litre of solvent.

The effect of temperature on solubility depends on the type of solute and solvent involved. This plays a crucial role in processes such as crystallisation, dissolution, precipitation, and solution preparation in various industries.

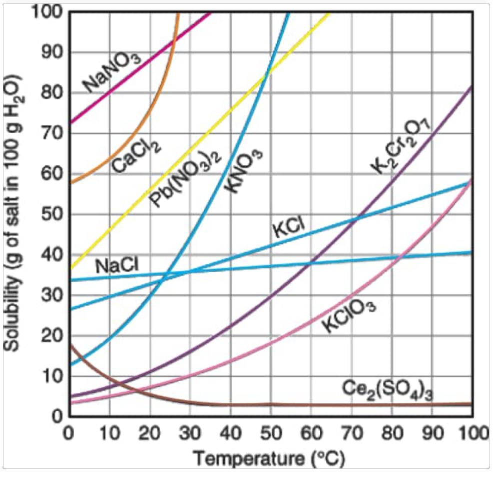
We observe that the solubility of tea leaves, coffee beans, and sugar increases with rising temperature and decreases with falling temperature. However, the solubility of certain substances, such as cerium sulphate and lithium carbonate, decreases with an increase in temperature.

Keep in Mind

Temperature affects not only the rate of dissolving but also the solubility. Generally, the rate at which a solute dissolves increases with a temperature rise and decreases when the temperature drops. However, the quantity of dissolved solute in the solution may either increase or decrease as the temperature of the solution rises.

2.7.1 Solubility of Solids in Liquids

Solids show strong attractive forces among their particles, requiring energy to overcome these forces when dissolved in a liquid. Generally, the solubility of solids in liquids increases with an increase in temperature. For example, the solubility of sugar (sucrose) in 100g of water at 0°C is 179g, whereas at 100°C it is 487g. The solubility of certain solids, such as NaCl, is slightly affected by an increase in temperature. The solubility of NaCl in 100g of water at 0°C is 35.7g, and at 100°C it is 39.2g. Conversely, the solubility of some solids decreases with increasing temperature. For example, the



Graph 2.1: Solublity of most compounds increase by increasing temperature

solubility of lithium carbonate at 0°C is 1.54g, and at 100°C it is approximately 0.72g. The solubility curves illustrating these relationships are shown in the figure, giving a graphical representation of the connection between temperature and the solubility of a solute in a specific solvent.

Table 2.4: Solubility in gram of solutes in 100g of water

Substance	Temperature (°C)					
	0	20	40	60	80	100
NaCl	35.7	35.9	36.4	37.1	38.0	39.2
$C_{12}H_{22}O_{11}$	179	204	238	287	362	487
Li_2CO_3	1.54	1.33	1.17	1.01	0.85	0.72
CO_2 at SP	0.335	0.169	0.0973	0.058	-	-
O_2 at SP	0.00694	0.00537	0.00308	0.00227	0.00138	0.00

2.7.2 Solubility of Gases in Liquids

The solubility of gases in water decreases with increasing temperature because higher temperatures elevate the kinetic energy of gas molecules, which in turn reduces the attractive forces with water molecules.

CHEMISTRY IN LIFE

Water and beverages contain dissolved gases such as oxygen, nitrogen, and carbon dioxide. When water is heated, bubbles appear before boiling as gases escape because the solubility of these gases decreases. The same principle explains why a soft drink loses its fizz quickly when warmed; dissolved CO_2 becomes less soluble and escapes more rapidly.

Interesting Information

Water used for various purposes contains dissolved gases, including nitrogen, oxygen, and carbon dioxide. When water is heated to make tea, small bubbles form inside a kettle before it boils, indicating the presence of dissolved gas molecules. As the water temperature rises, the solubility of dissolved gases decreases, causing them to escape from the water and form bubbles.

2.8 Formation of Unsaturated and Saturated Solutions

2.8.1 Unsaturated Solution

An unsaturated solution can dissolve more solute. In this case, the solvent has not reached the maximum solute concentration at a given temperature and pressure, allowing for more solute to dissolve.

2.8.2 Saturated Solution

A saturated solution is completely filled with solute and cannot dissolve more at a given temperature. If additional solute is added, it will remain undissolved and settle at the bottom.

2.8.3 Supersaturated Solution

A supersaturated solution contains more solute than is needed to form a saturated solution at a specific temperature. These solutions are unstable. The usual method is to first prepare a saturated solution at a high temperature, then cool it down. During cooling, some of the excess solute tends to crystallise, leaving behind a supersaturated solution. For example, a saturated solution of sodium thiosulphate in water at 20°C contains 20.9g of salt per 100cm³ of water. If we prepared a saturated solution at 30°C and cooled it to 20°C, it would contain more solute than 20.9g; it would be supersaturated.

ACITIVITY

Prepare a saturated sugar solution using hot water and allow it to cool undisturbed. Observe sugar crystals form; this process is similar to how rock candy is made.

2.8.4 Temperature and Saturation

As the temperature rises, many solutes dissolve more readily, leading to a saturated solution with higher concentrations of dissolved substances.

KEY POINTS

1. Matter is anything that has mass and occupies space.
2. Matter can be classified by its state (solid, liquid, gas, plasma) and by its composition (elements, compounds, and mixtures).
3. Solids have a definite shape and volume, high density, are not easily compressed, and their particles are tightly packed.
4. Liquids have a definite volume but no fixed shape, flow easily, have moderate density, and their particles are less tightly packed.
5. Gases have no definite shape or volume, have low density, are highly compressible, and their particles move freely.
6. Plasma is a high-energy state of matter found in stars and lightning.
7. Elements consist of one type of atom.
8. Compounds are formed when two or more elements combine chemically in a fixed ratio by mass.
9. Mixtures are formed by physically combining substances and can be homogeneous (uniform) or heterogeneous (not uniform).
10. The solubility of solids in liquids generally increases with temperature, while the solubility of gases in liquids decreases with temperature.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

I) The fundamental characteristic of matter is:

(a) colour (b) mass and volume
(c) temperature (d) pressure

ii) The following is **NOT** considered matter from a scientific perspective:
(a) water (b) sound
(c) trees (d) silver

iii) The state of matter characterised by having a definite shape and volume is:
(a) gas (b) liquid
(c) solid (d) plasma

iv) A typical property of liquids is:
(a) definite shape (b) definite structure
(c) high compressibility (d) ability to diffuse slowly

v) The average kinetic energy of gas molecules is directly proportional to:
(a) pressure of the gas (b) temperature of the gas
(c) volume of the gas (d) number of molecules

vi) The temperature at which a bose-Einstein condensate is typically formed is:
(a) room temperature (b) close to absolute zero
(c) boiling point (d) freezing point

vii) The typical size range of particles in a colloid is:
(a) 0.001nm to 0.01nm (b) 1nm to 1000nm
(c) 0.1nm to 1nm (d) 10nm to 100nm

viii) An example of an unsaturated solution is a solution:
(a) with dissolved solute at its maximum capacity
(b) with excess solute settled at the bottom
(c) that can dissolve more solute
(d) with equal amounts of solute and solvent

ix) Fog, classified as a type of mixture, is:
(a) solution (b) suspension
(c) colloid (d) compound

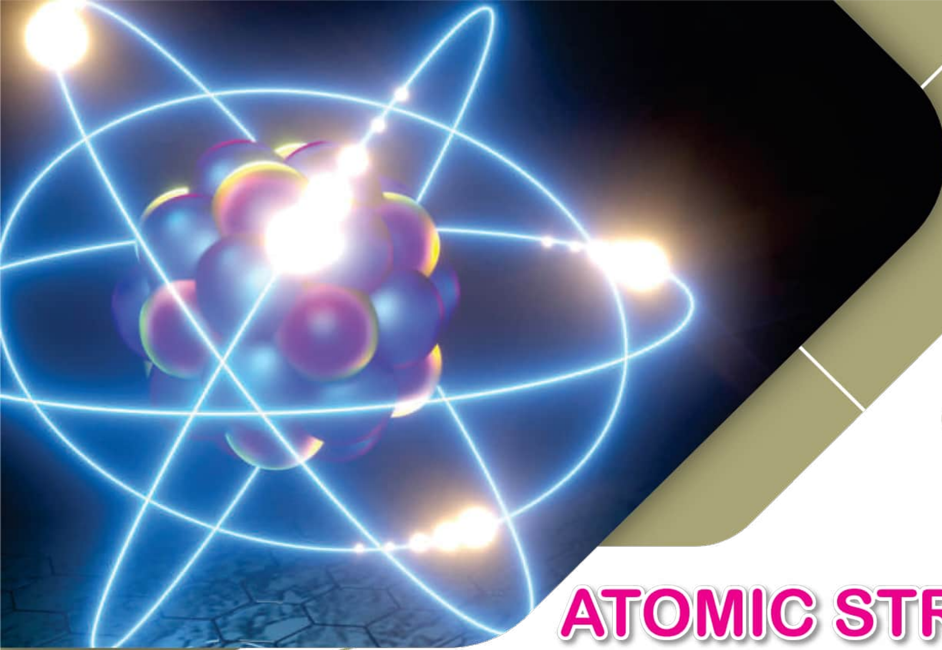
x) Solids are considered incompressible because:
(a) they have high kinetic energy (b) their particles are widely spaced
(c) their particles are tightly packed (d) they have a definite shape

B. Restricted Response Questions (RRQs)

i) What is diffusion? Provide one everyday example.
ii) Why are gases compressible, while solids are not?
iii) List two differences between homogeneous and heterogeneous mixtures.
iv) Why is milk considered a colloid instead of a solution?
v) Why do liquids flow, whereas solids do not?

C. Extended Response Questions (ERQs)

i) Differentiate solids, liquids, and gases based on their shape, volume, compressibility, and intermolecular forces.
ii) What is a solution? Describe the properties of a true solution and explain the terms solute and solvent with examples.
iii) Describe the effect of temperature on the solubility of solids and gases in liquids.
iv) Differentiate between element, compound, and mixture.
v) What are colloids? List their characteristics and provide four examples from daily life.



UNIT 3

ATOMIC STRUCTURE

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-08 Explain the structure of the atom as a central nucleus containing neutrons and protons surrounded by electrons in shells.

B-09 State that, orbits (shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus.

B-11 Explain that a nucleus is made up of protons and neutrons held together by strong nuclear force.

B-12 Explain that an atomic model is an aid to understand the structure of an atom.

B-13 State the relative charges and relative masses of subatomic particles (an electron, proton and neutron).

B-14 Interpret the relationship of masses and charges of subatomic particles.

B-15 Illustrate the path that positively and negatively charged particles would take under the influence of a uniform electric field.

B-16 Define proton number/atomic number as the number of protons in the nucleus of an atom.

B-17 Explain that the proton number is unique to each element and used to arrange elements in the Periodic Table.

B-19 Define nucleon number/ mass number as sum of number of protons and neutrons in the nucleus of an atom.

B-20 Define isotopes as different atoms of the same element that have same number of protons but different neutrons.

B-21 State that isotopes can affect molecular mass but not chemical properties of an atom.

B-22 Determine the number of protons and neutrons of different isotopes.

B-23 Define relative atomic mass as the average mass of isotopes of an element compared to $1/12^{th}$ of mass of an atom of Carbon-12.

B-25 Discuss the importance of isotopes using carbon dating and medical imaging as examples.

B-26 Describe the formation of positive (cation) and negative (anion) ions from atoms.

B-27 Interpret and use the symbols for atoms and ions.

B-28 Calculate relative atomic mass of an element from relative masses and abundance of isotopes.

B-29 Calculate the relative mass of an isotope given relative atomic mass and abundance of all stable isotopes.

Introduction

Everything in the universe is made up of tiny particles called atoms. These atoms are too small to see, yet they are the fundamental building blocks of everything around us, including the water we drink, the food we eat, and even our own bodies. Whether it's the trees, the mountains, the air we breathe, or the stars in the sky, everything consists of these little particles. This concept is known as the atomic theory of matter, which helps us understand how everything around us is connected.

Do You Know?

Atoms are so tiny that if you lined up 50 million hydrogen atoms in a row, they would only stretch across 1 millimetre.

3.1 Subatomic Particles of an Atom

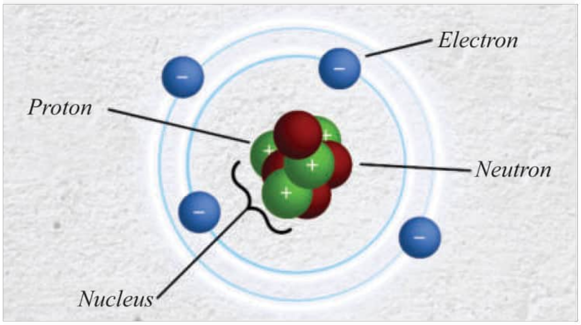


Fig. 3.1: Sub-atomic particles of beryllium

An atom is the fundamental building block of matter. It is the smallest unit of an element that retains all its chemical properties. Atoms consist of smaller particles known as subatomic particles. The primary subatomic particles include protons, neutrons, and electrons. Protons and neutrons are collectively referred to as nucleons. Protons have a positive charge, neutrons have no charge (they are neutral), and electrons carry a negative charge. Protons and neutrons are located in the nucleus of an atom, while electrons orbit around the nucleus in shells. Protons and neutrons have almost identical masses, whereas electrons have a much smaller mass. A proton has a charge equal in magnitude and opposite in sign to that of an electron. Atoms are electrically neutral, meaning the number of electrons equals the number of protons in the nucleus. The mass of an atom is determined mainly by the number of protons and neutrons it contains. The mass of electrons is neglected in calculations of the mass of an atom.

Table 3.1: The key characteristics of electrons, protons, and neutrons

Subatomic Particle	Symbol	Relative Charge	Relative Mass
Electron	e^-	−1	0.00054
Proton	p^+	+1	1.007276
Neutron	n°	0	1.008665

Self-Assessment

A diamond is a precious non-metallic element represented by the symbol $^{12}_{6}C$. Complete the statement by filling in the blanks:

At the centre of a diamond atom lies the ______, which contains **6** ______ and **6** ______.

3.2 Energy of Electrons in Orbits

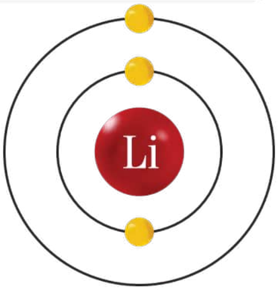


Fig. 3.2: Lithium atom

The electrons in an atom are arranged in orbits, also known as shells or energy levels (in modern concepts), surrounding the nucleus. The energy of an electron is determined by its position relative to the nucleus. Electrons closer to the nucleus experience a stronger force of attraction and therefore have lower energy, while those farther away experience a weaker force of attraction and possess higher energy. For example, lithium has an atomic number of 3, with two electrons in the first shell and one electron in the second shell. Consequently, the electron in the second shell has higher energy compared to those in the first shell and greater average distance from the nucleus.

AMAZING FACT

The orbit was circular and two-dimensional, whereas the shell is spherical and three-dimensional (3D).

3.3 Nuclear Force

The fundamental force known as the nuclear force binds protons and neutrons together within the nucleus of an atom. This force overcomes the electrostatic repulsion that occurs naturally between positively charged protons. The strong nuclear force operates over very short distances, 0.8 to 2.5 femtometres ($1 fm = 10^{-15} m$). It is significantly stronger than the electromagnetic force, which causes protons to repel each other due to their positive charges. When nucleons are within this range, the strong force holds them closely together.

AMAZING FACT

More than 99.9% of the mass of an atom is in the nucleus, yet the nucleus itself occupies only a tiny fraction of the total volume of the atom.

3.4 Atomic Models

Atoms are so small that they cannot be seen with the naked eye. Scientists use models to help understand their behaviour and visualise their appearance. Over time, different atomic models have been developed, each building on previous discoveries. Let us take a brief journey through some key models in history.

FUN FACT

The word "atom" originates from the Greek term "atomos," meaning indivisible, because early philosophers believed atoms could not be divided.

3.4.1 Thomson's Plum Pudding Model (1904)

According to Thomson, an atom is a sphere of positive charge with negatively charged particles (electrons) embedded within it. This confirms the existence of electrons inside the atom and explains the electrical neutrality of an atom.

3.4.2 Rutherford's Atomic Model (1911)

Rutherford's gold foil experiment showed that an atom has a tiny, positively charged nucleus at its centre, surrounded by negatively charged electrons moving around it. The nucleus is much smaller than the atom, and the entire mass of an atom is concentrated within the nucleus. It cannot explain the stability and spectrum of an atom.

3.4.3 Bohr's Atomic Model (1913)

Bohr's model explained the stability of atoms and accounted for atomic spectra and the ionisation of gases. According to this model, an atom consists of a positively charged nucleus and electrons revolving around it in fixed-energy circular orbits. They lose or gain energy only when they jump from one orbit to another.

3.4.4 Quantum Mechanical Model (1927)

The quantum model uses quantum mechanics to describe the behaviour of electrons within atoms. This model illustrates the arrangement of electrons in subshells and orbitals, challenging the concept of fixed orbits.

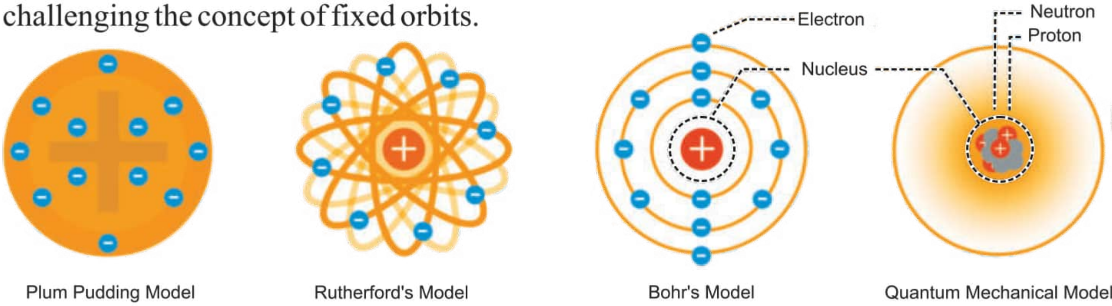


Fig. 3.3: Atomic models

Real-world connections

The development of atomic models established the basis for technologies such as lasers, nuclear power, MRI machines, and microchips.

3.5 Behaviour of Charged Particles in a Uniform Electric Field

The entire world consists of charged particles, and it is the interaction between these particles that allows us to observe all the various phenomena occurring in our universe. When a charged particle generates a force that affects another nearby charged particle, this phenomenon is known as an electric field. If the field has a consistent strength and direction, it is referred to as a uniform electric field. Now, let us discuss how protons and electrons behave in a uniform electric field. Protons and electrons carry equal but opposite charges and forces. Therefore, when placed in a uniform electric field, the forces acting on them are equal in magnitude.

3.5.1 Positively Charged Particles (Protons)

Protons have a positive charge and will move toward the negative plates in an electric field.

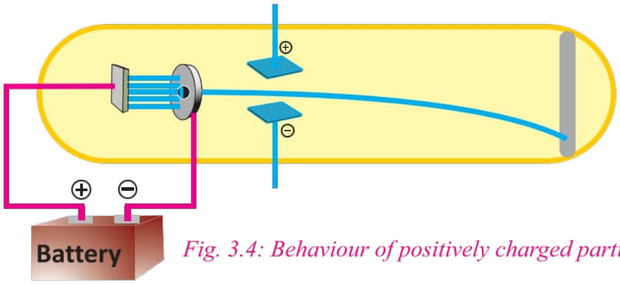


Fig. 3.4: Behaviour of positively charged particles

3.5.2 Negatively Charged Particles (Electrons)

Electrons have a negative charge and will move toward positive plates in an electric field.

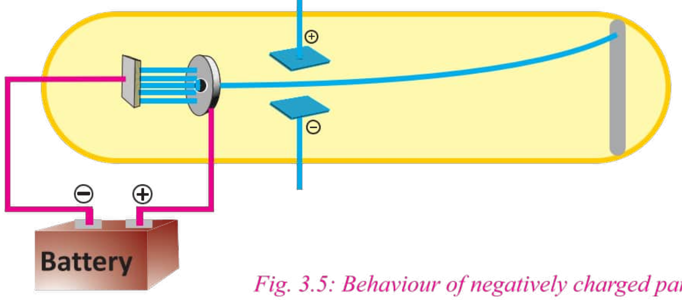


Fig. 3.5: Behaviour of negatively charged particles

The behaviour of a charged particle in an electric field is very important because electric fields are used in various applications, such as particle accelerators.

3.6 Atomic Number

The number of protons in the nucleus of an atom is known as the atomic number (proton number) and is represented by the symbol Z. Each element has a unique atomic number, and no two elements can have the same atomic number. This uniqueness is essential for classifying and identifying elements. For example, a nitrogen atom has seven protons; if you were to add another proton to the nitrogen nucleus, it would no longer be a nitrogen atom; instead, it would transform into an oxygen atom with eight protons.

The elements in the periodic table are arranged in order of increasing atomic number, which helps predict their behaviour and properties.

Real-World Connections

The atomic number determines the position of an element in the periodic table, which in turn affects modern industries from steel manufacturing to semiconductors.

Keep in Mind:

The elements can be identified by the number of protons (atomic number) in their atoms. Atoms that have the same number of protons belong to the same element, while atoms with different numbers of protons are from different elements.

3.7 Mass Number

The mass number of an element is the total number of protons and neutrons in the nucleus of an atom, denoted by the symbol A. It is also known as the nucleon number.

Mass Number = number of protons + number of neutrons

For example, oxygen-16 has a mass number of 16, which means it contains eight protons and eight neutrons; hence, its mass number is 16.

The atomic number (Z) and mass number (A) of an element are usually expressed as follows:

$$^{A}_{Z}X$$

Where X is the symbol of an element, A is the mass number, and Z is the atomic number. For example, nitrogen-14 is written as $^{14}_{7}N$, where N is the symbol for nitrogen, 14 is the mass number, and 7 is the atomic number.

Keep in Mind:

In a neutral atom, the number of electrons equals the number of protons; however, the number of neutrons in the nucleus is not always the same as the number of protons.

Table 3.2: Number of protons, electrons and neutrons in some elements

Element	Symbol	Atomic Number (Z)	Mass Number (A)	Number of Protons	Number of Electrons	Number of Neutrons
Hydrogen	H	1	1	1	1	0
Carbon	C	6	12	6	6	6
Neon	Ne	10	20	10	10	10
Sulphur	S	16	32	16	16	16
Potassium	K	19	39	19	19	20
Zinc	Zn	30	65	30	30	35

3.8 Isotopes

Atoms of the same element have the same number of protons but different mass numbers. They show similar chemical properties due to their identical electronic structures. However, they have different physical properties due to their different masses (or number of neutrons). For example, hydrogen has three isotopes: protium (^1_1H), deuterium (^2_1H) and tritium (^3_1H), while chlorine has two isotopes: chlorine-35 ($^{35}_{17}Cl$) and chlorine-37 ($^{37}_{17}Cl$). Isotopes are usually named using the name of the element followed by their mass number.

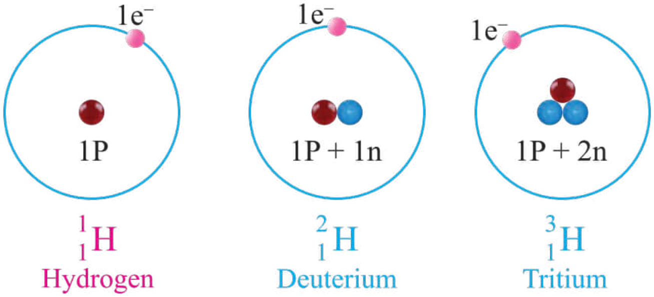


Fig. 3.6: Isotopes of hydrogen

Keep in Mind:

The isotopes of an element interact with other elements in the same way because they have the same number of electrons. The difference in isotopic mass does not significantly affect the chemical behaviour of an element.

Example 3.1 **Calculate the number number of protons and neutrons present in the nucleus of:**
(i) carbon-14 ($^{14}_{6}C$), (ii) fluorine-19 ($^{19}_{9}F$) and (iii) argon-40 ($^{48}_{18}Ar$).

SOLUTION

i) Carbon-12:

The atomic number of carbon is 6, so it has 6 protons. Carbon-14 has an atomic mass of 14, so the number of neutrons (A–Z) = 14 – 6 = 8.

ii) Fluorine-19:

The atomic number of fluorine is 9, so it has 9 protons. Fluorine-19 has an atomic mass of 19, so the number of neutrons (A–Z) = 19 – 9 = 10.

iii) Argon-40:

The atomic number of argon is 18, so it has 6 protons. Argon-18 has an atomic mass of 18, so the number of neutrons (A–Z) = 40 – 18 = 22.

Test Yourself: Calculate the number of protons, electrons and neutrons in each of $^{28}_{14}Si$, $^{29}_{14}Si$, and $^{30}_{14}Si$.

3.8.1 Importance of Isotopes

Isotopes play an essential role in various scientific disciplines and have numerous applications, including carbon dating and medical imaging. This discussion will explore the significance of isotopes in these two particular fields.

Carbon Dating

Carbon-14, one of three carbon isotopes, is crucial for carbon dating, which determines the age of artefacts and organic materials by measuring their decay. With a half-life of 5730 years, it is effective for dating materials up to 50,000 years old. After an organism's death, comparing the remaining Carbon-14 to the expected amount reveals the time of death.

Interesting Information

Imagine finding something ancient, like wood or a bone. Scientists estimate its age by measuring remaining carbon-14; less carbon-14 indicates an older object.

Medical Imaging

Medical imaging plays a crucial role in diagnosing and treating various medical conditions. It employs radioactive isotopes that emit positrons or gamma rays. Isotopes like Fluorine-18 in Positron Emission Tomography (PET) scans help in the detection of diseases. Technetium-99m is used in Single Photon Emission Computed Tomography (SPECT) to assess organ function, while Iodine-131 is used to image and treat thyroid disorders. Thallium-201 aids in cardiac imaging and tracing urine flow, helping to detect kidney issues. These isotopes, when attached to molecules, serve as tracers, facilitating accurate imaging through various techniques, including X-rays, Magnetic Resonance Imaging (MRI), Computed Tomography (CT) scans, and ultrasound imaging.

CHEMISTRY IN LIFE

PET and SPECT use radioactive isotopes to assess organ function, with PET being highly sensitive for cancer and brain studies, and SPECT widely used in cardiac, bone, and thyroid scans. X-rays quickly image bones, the chest, and teeth. Ultrasound provides safe, real-time images of organs, and blood flow via sound waves. MRI uses magnets and radio waves to produce detailed images of the brain, spine, joints, tumours, and soft tissues.


PET


SPECT


X-Ray


Ultrasound


MRI

3.9 Relative Atomic Mass

The relative atomic mass of an element is the average mass of its isolopes compared to $1/12^{th}$ of the mass of a carbon-12 atom. This value is based on the relative abundance of each isotope of the element and is used to compare the masses of different elements. The unit used to express atomic and molecular masses is the atomic mass unit (amu). One amu is defined as 1/12 of the mass of a single carbon-12 atom, which is approximately 1.66×10^{-27} kilograms.

$$\text{Relative Atomic Mass}(A_r) = \frac{\Sigma(\text{Mass of each isotope} \times \text{Percentage abundance of each isotope})}{100}$$

The formula for calculating the relative atomic mass of an element with multiple isotopes is:

$$\text{Relative Atomic Mass} = \frac{(M_1 \times A_1) + (M_2 \times A_2) + \ldots + (M_n \times A_n)}{100}$$

Where:

- $M_1, M_2, \ldots, M_n$ = Masses of the respective isotopes.
- $A_1, A_2, \ldots, A_n$ = Percentage abundances of the respective isotopes.

For example, a carbon sample consists of a mixture of two isotopes, carbon-12 and carbon-13. These isotopes occur in specific proportions in the sample: 89.89% carbon-12 and 1.11% carbon-13. The relative atomic mass of carbon in the sample can be calculated as:

$$\text{Relative Atomic Mass of Carbon} = \frac{(12 \times 98.89) + (13 \times 1.11)}{100}$$

$$= \frac{1186.68 + 14.43}{100} = \frac{1201.11}{100} = 12.01$$

Therefore, the relative atomic mass of carbon is approximately 12.01amu.

3.9.2 Calculation of the Relative Mass of an Isotope

The formula to calculate the relative atomic mass of an element is:

$$\text{Relative Atomic Mass} = \frac{(M_1 \times A_1) + (M_2 \times A_2) + \ldots + (M_n \times A_n)}{100}$$

If you want to calculate the relative mass of a specific isotope, such as Isotope 1, the formula is:

$$M_1 = \frac{A_r \times 100 - (M_2 \times A_2 + M_3 \times A_3)}{A_1}$$

Consider the example of chlorine, which has two stable isotopes. The relative atomic mass of chlorine is 35.5, the mass of Isotope 2 is 35, and the relative abundance of Isotope 2 is 75%. The relative abundance of Isotope 1 is 25%, and we want to calculate the mass of Isotope 1 as follows:

$$M_1 = \frac{35.5 \times 100 - (35 \times 75)}{25} = \frac{3550 - 2625}{25} = \frac{925}{25} = 37\text{amu}$$

Hence, the mass of Isotope 1 is 37amu.

Test Yourself: Chlorine contains two stable isotopes, and its relative atomic mass is 35.5. Isotope A has a mass of 37 and is present in natural abundance at 25%. Determine the mass of isotope B, knowing that its relative abundance is 75%.

3.10 Symbols of Elements, and Ions

3.10.1 Symbols of Elements

Each element is represented by a unique standard symbol consisting of one or two letters. The first letter is always capitalised, while the second letter, if present, is written in lowercase. For example, the symbols for beryllium and calcium are Be and Ca, respectively. Some symbols are derived from Latin names of elements; for example, the symbol for sodium (natrium) is Na, potassium (kalium) is K, silver (argentum) is Ag, gold (aurum) is Au, and lead (plumbum) is Pb.

3.10.2 Symbols of Ions

An ion is an atom or group of atoms that carries a charge due to the loss or gain of electrons. The symbol of an ion is written with the symbol of the element followed by the charge as a superscript. For example, the symbols for cations are Na^+, Ca^{2+}, Fe^{3+}, and for anions are Cl^-, O^{2-}, N^{3-}. Polyatomic ions such as SO_4^{2-}, NO_3^-, and NH_4^+ are also written in this way.

Table. 3.3: Symbols of some common elements and ions

Element	Symbol	Latin Name (if applicable)	Common Ion(s)
Hydrogen	H	-	H^+
Helium	He	-	–
Lithium	Li	-	Li^+
Beryllium	Be	-	Be^{2+}
Carbon	C	-	–
Nitrogen	N	-	N^{3-}
Sodium	Na	Natrium	Na^+
Aluminium	Al	-	Al^{3+}
Potassium	K	Kalium	K^+
Calcium	Ca	-	Ca^{2+}
Iron	Fe	Ferrum	Fe^{2+}, Fe^{3+}
Copper	Cu	Cuprum	Cu^+, Cu^{2+}
Silver	Ag	Argentum	Ag^+
Tin	Sn	Stannum	Sn^{2+}, Sn^{4+}
Gold	Au	Aurum	Au^+, Au^{3+}
Mercury	Hg	Hydrargyrum	Hg^+, Hg^{2+}
Lead	Pb	Plumbum	Pb^{2+}, Pb^{4+}

3.11 Formation of Ions

An ion is a charged particle formed when an atom or molecule gains or loses electrons. Ions form when an atom or molecule has an unequal number of protons and electrons. If it has more electrons than protons, it carries a negative charge and is called an anion. Conversely, if it has more protons than electrons, it carries a positive charge and is known as a cation.

3.11.1 Formation of Cations

Cations are positively charged ions formed when an atom loses one or more electrons. Atoms of metals such as sodium, calcium, and aluminium commonly form cations. For example, a sodium atom (Na) has 11 protons and 11 electrons in its neutral form, with one electron in its outermost shell (2,8,1). When it loses the outer electron, it becomes a sodium cation (Na^+). This Na^+ ion now has 11 protons and 10 electrons, resulting in a net positive charge of 1+.

$$\underset{2,8,1}{Na_{(g)}} \longrightarrow \underset{2,8}{Na^+_{(g)}} + 1e^-$$

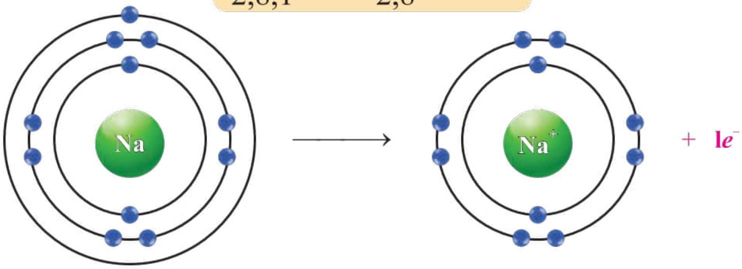


Table 3.4: Some important cations

Common Name	Symbol	Common Name	Symbol
Aluminum	Al^{3+}	Nickel	Ni^{2+}
Calcium	Ca^{2+}	Magnesium	Mg^{2+}
Chromium(II) (Chromous)	Cr^{2+}	Chromium(III) (Chromic)	Cr^{3+}
Copper(I) (Cuprous)	Cu^{+}	Copper(II) (Cupric)	Cu^{2+}
Lead(II) (Plumbous)	Pb^{2+}	Lead (IV) (Plumbic)	Pb^{4+}
Silver	Ag^{+}	Potassium	K^{+}
Hydrogen	H^{+}	Sodium	Na^{+}
Iron(II) (Ferrous)	Fe^{2+}	Iron(III) of Frrric	Fe^{3+}

3.11.2 Formation of Anions

Anions are negatively charged ions formed when an atom gains one or more electrons. Non-metal atoms, such as nitrogen, oxygen, and chlorine, commonly form anions. For example, a chlorine atom (Cl) has seven electrons in its outermost shell (2,8,7) and needs one more electron to complete its octet. When it gains this electron, it becomes a chloride anion (Cl^-). This Cl^- ion now has 17 protons and 18 electrons, resulting in a net negative charge of 1–.

$$\underset{2,8,7}{Cl_{(g)}} + 1e^- \longrightarrow \underset{2,8,8}{Cl^-_{(g)}}$$

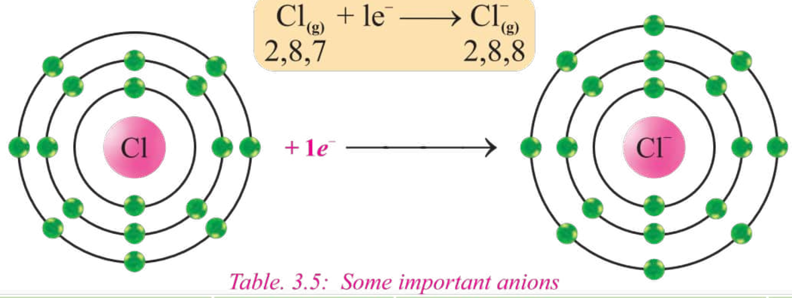


Table. 3.5: Some important anions

Common Name	Symbol	Common Name	Symbol
Bromide	Br^-	Iodide	I^-
Chloride	Cl^-	Nitride	N^{3-}
Fluoride	F^-	Oxide	O^{2-}
Hydride	H^-	Sulphide	S^{2-}

KEY POINTS

1. The atomic theory of matter states that all substances are made up of indivisible particles called atoms, the basic building blocks of matter.
2. Atoms are made up of subatomic particles: protons and neutrons in the nucleus, and electrons revolving around the nucleus at specific energy levels.
3. Protons carry a positive charge, neutrons are neutral, and electrons are negatively charged; protons and neutrons have nearly equal mass, while electrons have negligible mass.
4. A strong nuclear force holds protons and neutrons together in the nucleus, counteracting electrostatic repulsion between protons.
5. In an electric field, protons move towards the negative plate, while electrons are attracted to the positive plate.
6. The atomic number (Z) is equal to the number of protons in the nucleus and identifies the element.
7. The mass number (A) is the sum of protons and neutrons in the nucleus and represents the mass of the atom.
8. Isotopes are atoms of the same element with the same atomic number but different mass numbers; they have identical chemical properties but different physical properties.
9. The relative atomic mass of an element is the average mass of its atoms compared to one-twelfth of the mass of a carbon-12 atom, taking into account isotopic abundances.
10. Atoms form ions by losing or gaining electrons: loss results in positively charged cations, while gain results in negatively charged anions.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The atomic number (Z) of an element is defined as the number of:
- (a) neutrons in the nucleus
- (b) electrons in all shells
- (c) protons in the nucleus
- (d) protons and neutrons together

ii) The relative atomic mass of an element is based on:
- (a) mass of the lightest isotope of the element
- (b) average mass of isotopes compared with 1/12th mass of a carbon-12 atom
- (c) total number of protons and neutrons in the nucleus
- (d) mass of a single atom measured in kilograms

iii) The charge of a proton is:
- (a) negative
- (b) positive
- (c) neutral
- (d) variable

iv) The atomic model in which electrons are arranged in specific orbits or shells is:
- (a) thomson's plum pudding model
- (b) rutherford's atomic model
- (c) bohr's atomic model
- (d) quantum mechanical model

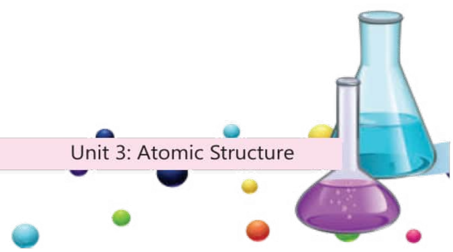
v) The force responsible for binding protons and neutrons together in the nucleus is:
- (a) electromagnetic force
- (b) gravitational force
- (c) strong nuclear force
- (d) weak nuclear force

vi) A neutron has a charge that is:
- (a) positive
- (b) negative
- (c) neutral
- (d) variable

vii) The atomic number of an element is equal to the number of:
- (a) protons
- (b) neutrons
- (c) electrons
- (d) nucleons

viii) The mass of a proton compared to that of an electron is:
- (a) 1/1836
- (b) 1836
- (c) 1839
- (d) 1842

ix) An electron carries a charge equal in magnitude to that of a proton but opposite in sign. Its charge is:
- (a) +1
- (b) –1
- (c) 0
- (d) 1/1837



x) The mass number of an element is:

(a) Number of protons (b) Number of neutrons

(c) Sum of protons and electrons (d) Sum of protons and neutrons

B. Restricted Response Questions (RRQs)

i) What does relative atomic mass mean?

ii) Why do atoms have no overall charge?

iii) Briefly explain why a sodium ion has a charge of 1+ while a chloride ion has a charge of 1–.

iv) Why do scientists conclude that electrons revolve around the nucleus and not the neutron?

v) How does the mass of a proton compare to that of an electron?

vi) Some chemical elements have symbols derived from their Latin names. Identify the element for each given Latin name below:

a) Natrium b) Kalium

c) Ferrum d) Plumbum

f) Aurum

Provide the modern English name and the chemical symbol for each.

C. Extended Response Questions (ERQs)

i) Explain the structure of the atom, including its nucleus and electron shells.

ii) What are isotopes, and how do they differ in terms of protons and neutrons?

iii) What is relative atomic mass, and how is it calculated from isotopes?

iv) How can isotopes be used in practical applications such as carbon dating and medical imaging?

D. Project

Construct a simple 3D model of a carbon atom using coloured balls to represent protons, neutrons, and electrons, and label the nucleus and electron shells.

UNIT
4

CHEMICAL BONDING

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-30 Describe that noble gas electronic configuration, octet and duplet rules help predict chemical properties of main group elements.

B-31 Compare between the formation of cations and anions.

B-32 Account for the electropositive and electronegative nature of metals and non-metals.

B-33 Define ionic, covalent, coordinate covalent and metallic bonds.

B-34 Differentiate between ionic compounds and covalent compounds.

(The following points need 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 bond as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations.)

B-35 Explain the properties of compounds in terms of bonding and structure.

B-36 Compare uses and properties of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.

B-37 Interpret the strength of forces of attraction and their impact on melting and boiling points of ionic and covalent compounds.

B-38 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.

B-39 Recognise that some substances can ionise when dissolved in water (e.g., acids dissolve in water and conduct electricity).

B-40 Justify the suitability of usage of graphite, diamond and metals for industrial purposes, (Some examples may include:

a) graphite as lubricant or an electrode

b) diamond in cutting tools

c) metals for wires, and sheets)

B-41 Draw the structure of ionic and covalent compounds along with their formation. (Some examples can include:

a) Ionic bonds in binary compounds such as NaBr, NaF, and CaCl using dot-and-cross diagrams and Lewis-dot structures.

b) covalent bonds in 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-dot structures.)

Introduction

Everything around us is made up of matter. Matter is composed of atoms that combine to form molecules, which can exist in different states. The forces that bind atoms together in a molecule are known as chemical bonds or chemical forces. These bonding forces, which hold the atoms in position, form the foundation of this unit.

4.1 Noble Gas Configuration and the Octet–Duplet Rules in Predicting Chemical Properties

Noble gases such as helium, neon, and argon are characterised by completely filled outermost shells, which makes them chemically stable and largely unreactive. Helium has two electrons ($1s^2$), while the other noble gases possess eight valence electrons (ns^2np^6). This complete valence shell explains their notable inertness.

Do You Know?

Neon lights glow due to the inert nature of noble gases. When electricity passes through a glass tube filled with neon, the stable electronic configuration prevents chemical reactions but excites electrons enough to produce a bright red-orange glow.

In 1916, G.N. Lewis introduced the **Octet Rule**, which states that atoms tend to gain, lose, or share electrons to complete their valence shell with eight electrons, achieving the stable electronic configuration of the nearest noble gas. Similarly, very light elements, such as hydrogen, lithium, and beryllium, follow the **Duplet Rule**, aiming for two electrons in their valence shell, similar to helium.

Groups IA and IIA elements tend to lose one or two electrons, respectively, to achieve the electron configuration of the previous noble gas. For example, sodium (Na) loses one electron, and magnesium (Mg) loses two electrons to attain the configuration of neon.

Groups VA–VIIA elements tend to gain three, two, or one electron(s), respectively, to complete their octet. For example, Nitrogen (N) gains three electrons to achieve the configuration of neon.

Valence electrons thus determine whether an element will form ionic bonds (by electron transfer) or covalent bonds (by electron sharing) to achieve noble gas stability.

Table 4.1: Electronic configuration of noble gases

Noble Gases	Symbols	Atomic Number	Shell Electron Configuration	No. of Electrons in the Valence Shell
Helium	He	2	2	2
Neon	Ne	10	2, 8	8
Argon	Ar	18	2, 8, 8	8
Krypton	Kr	36	2, 8, 18, 8	8
Xenon	Xe	54	2, 8, 18, 18, 8	8
Radon	Rn	86	2, 8, 18, 32, 18, 8	8

Predicting Chemical Properties

By comparing the electronic configurations of representative elements with those of noble gases, chemists can predict their reactivity. Elements with noble gas-like configurations are generally stable and show very low reactivity. In contrast, elements with incomplete configurations are highly reactive and tend to form bonds to achieve stability.

Keep in Mind

Representative elements are located in the periodic table in groups IA to VIIA. They have incomplete outer electron shells, which makes them chemically reactive. In contrast, noble gases (group VIIIA) have fully filled outer electron shells, making them chemically inert and not classified as representative elements.

4.2 Formation of Cations and Anions

The concept of ions was introduced in the previous chapter (atomic structure), where we discussed how atoms lose or gain electrons to achieve stable noble gas configurations. To briefly recall, cations are positively charged ions formed when atoms (usually metals) lose electrons, while anions are negatively charged ions formed when atoms (usually nonmetals) gain electrons. In this chapter, we will further explore how these ions participate in bond formation, particularly in ionic compounds.

Self-Assessment

1. Describe how a sodium atom forms a cation, including its electronic configuration before and after ion formation.
2. Explain how a chlorine atom forms an anion, and show its electronic configuration before and after ion formation.

Table 4.2: Comparison between cations and anions

Feature	Cations	Anions
Definition	Formed by the loss of one or more electrons from an atom.	Formed by the gain of one or more electrons from an atom.
Charge	They are positively charge.	They are negatively charge.
Size	Smaller than the parent atom.	Larger than the parent atoms.
Electronegativity	Associated with low electronegativity.	Associated with high electronegativity.
Ionic compounds	Combine with anions to form ionic compounds.	Combine with cations to form ionic compounds.
Electron configuration	Attained a stable electronic configuration of noble gases by losing electrons.	Attained a stable electronic configuration of noble gases by gaining electrons.
Examples	H^+, Mg^{2+}, Al^{3+}	F^-, O^{2-}, N^{3-}

CHEMISTRY IN LIFE

Sodium (Na^+) and potassium (K^+) ions regulate heartbeat and nerve transmission. Without the proper balance of these ions, our muscles would not contract properly.

4.3 Electropositive and Electronegative Nature of Metals and Nonmetals

Electropositivity and electronegativity are essential characteristics of chemical elements. These properties are related to their ability to gain or lose electrons during chemical reactions and their positions within the periodic table.

4.3.1 Electropositive Nature of Metals

Metals show an electropositive nature, meaning they tend to lose electrons easily in chemical reactions to form positive ions. This is because metals have low ionisation energies, making it easier for them to lose electrons and achieve a stable electronic configuration. Metals are located on the left side of the periodic table, where they have fewer valence electrons and a greater tendency to lose electrons. For example, consider calcium (Ca) in Group 2. It has two electrons in its outermost shell. To achieve a stable electron configuration similar to that of argon, calcium readily loses these electrons to form a Ca^{2+} cation:

$$\underset{2,8,8,2}{Ca} \longrightarrow \underset{2,8,8}{Ca^{2+}} + 2e^-$$

4.3.2 Electronegative Nature of Nonmetals

Nonmetals, on the other hand, possess an electronegative nature and tend to gain electrons easily in chemical reactions, forming negative ions. They have high ionisation energies, making it more difficult for them to lose electrons. They are located in the top right side of the periodic table, where they have more valence electrons and a greater tendency to gain electrons. For example, consider oxygen (O) in Group VIA, which has six electrons in its outermost shell. By gaining two electrons, oxygen achieves a stable electron configuration similar to that of neon:

$$\underset{2,6}{O} + 2e^- \longrightarrow \underset{2,8}{O^{2-}}$$

Chemical bonds between atoms are typically classified as either ionic or covalent.

Chemistry in Your Body

The electronegativity of oxygen gives water its polarity. This enables water to dissolve salts and nutrients, allowing blood to carry essential substances throughout the body.

4.4 Ionic, Covalent and Metallic Bonds

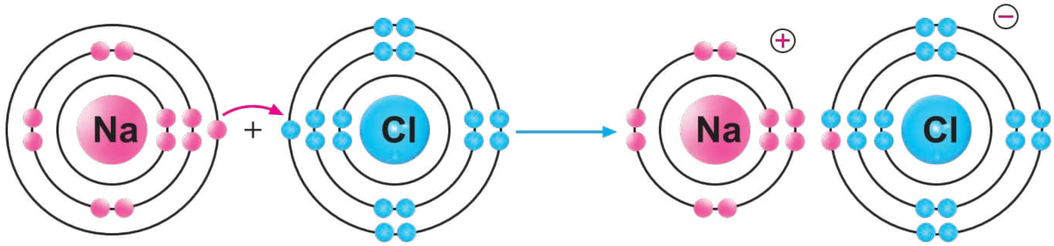
Chemical bonds are the forces that hold atoms together to form molecules and compounds. Interactions between atoms create chemical bonds through sharing or transferring electrons to reach a stable electron configuration.

4.4.1 Ionic Bond

Ionic bonds form through a strong electrostatic force of attraction between oppositely charged ions. When an atom with high electronegativity (such as a non-metal) bonds with an atom with low electronegativity (such as a metal), the non-metal atom attracts electrons away from the metal atom. This results in the formation of positively charged cations and negatively charged anions.

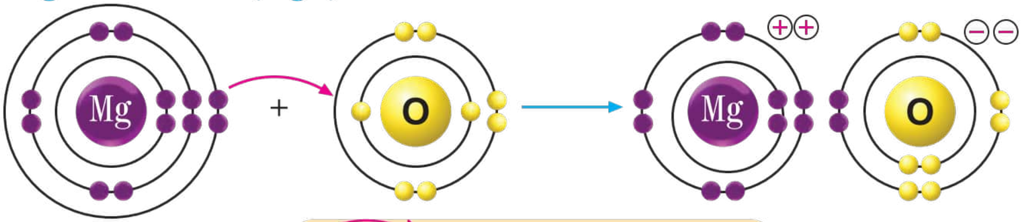
Examples of ionic bonds include:

1. Sodium Chloride (NaCl)



$$\dot{Na} + \cdot\ddot{\underset{\cdot\cdot}{Cl}}\!: \longrightarrow Na^+ :\ddot{\underset{\cdot\cdot}{Cl}}\!:^-$$

2. Magnesium Oxide (MgO)



$$Mg + :\ddot{O}: \longrightarrow [Mg]^{2+}\ [:\ddot{O}:]^{2-}$$

4.4.2 Covalent Bond

Covalent bonds are strong attractions formed by the mutual sharing of electrons between two atoms. This bond forms when there is a small difference in electronegativity between the atoms involved. Covalent bonds can be either nonpolar (equal sharing) or polar (unequal sharing), depending on the differences in electronegativity.

1. Types of Covalent Bond

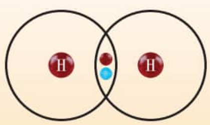
There are three types of covalent bonds: single covalent bonds, double covalent bonds, and triple covalent bonds.

Single Covalent Bond

The bond formed by the sharing of one electron pair is called a single covalent bond. It is represented by a single dash (–).

Example 1

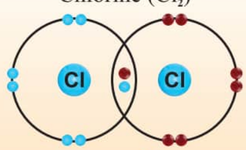
Hydrogen (H_2)



H—H

Example 2

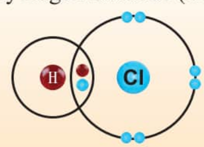
Chlorine (Cl_2)



$:\ddot{Cl}—\ddot{Cl}:$

Example 3

Hydrogen Chloride (HCl)



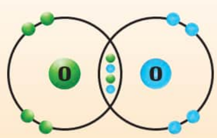
$H—\ddot{Cl}:$

Double Covalent Bond

The bond formed by the sharing of two electron pairs is called a double covalent bond. It is represented by a double dash (=).

Example 1

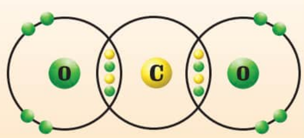
Oxygen (O_2)



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

Example 2

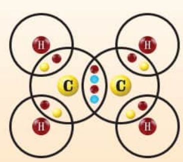
Carbon Dioxide (CO_2)



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

Example 3

Ethene (C_2H_4)



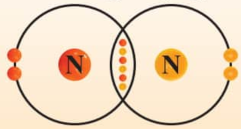
$CH_2=CH_2$

Triple Covalent Bond

The bond formed by the sharing of three electron pairs is called a triple covalent bond. It is represented by a triple dash (≡).

Example 1

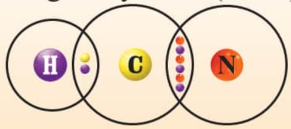
Nitrogen (N_2)



:N≡N:

Example 2

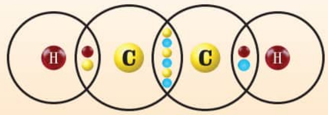
Hydrogen Cyanide (HCN)



H—C≡N:

Example 3

Ethyne (C_2H_2)



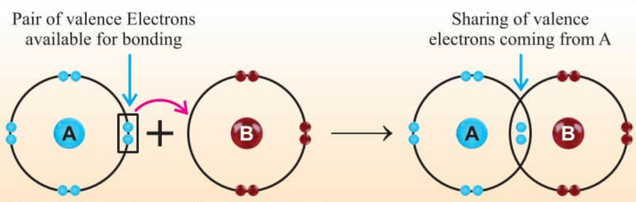
H—C≡C—H

Keep in Mind

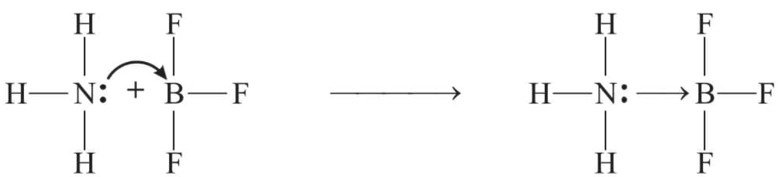
Ionic compounds have high melting and boiling points due to the strong electrostatic forces between ions in their crystal lattice. Sufficient heat energy is required to break the ionic bonds and separate the ions during phase transitions. Conversely, covalent compounds have weaker forces, resulting in lower melting and boiling points. The sharing of electrons between atoms in covalent compounds can be more easily disrupted by heat.

4.4.3 Coordinate Covalent Bond (Dative Bond)

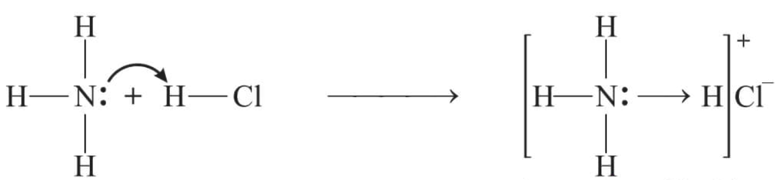
A coordinate covalent bond is formed when a shared pair of electrons is donated by one of the bonded atoms. An atom that donates a shared pair of electrons is called a donor, while an atom that accepts a shared pair of electrons is called an acceptor. This bond is shown by an arrow (→).



Other examples are:



Ammonia Boron Triflouride

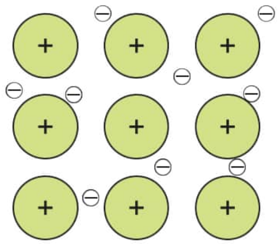


Ammonium chloride

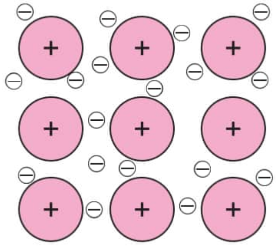
Interesting Information The hydronium ion (H_3O^+) forms whenever an acid dissolves in water. This ion is the actual carrier of acidity in solutions.

4.4.4 Metallic Bond

Metallic bonds are formed by a strong electrostatic attraction between a cloud (or sea) of delocalised electrons and positively charged metal ions. The delocalised electrons can move freely within the metal lattice, where they are attracted to all the cations. In this bond, metal atoms lose their valence electrons, forming a sea of electrons. Positively charged metal ions are then attracted to the delocalised electrons.



Sodium metal



Magnesium metal

Fig. 4.1: Metallic bonds in sodium and magnesium metals

Table 4.3: Differences between ionic, covalent and metallic bonds

Feature	Ionic Bonds	Covalent Bonds	Metallic Bonds
Definition	Strong electrostatic attraction between oppositely charged ions.	Strong electrostatic attraction between shared electrons and two nuclei.	Strong electrostatic attraction between a cloud of delocalised electrons and positively charged cations.
Bond Strength	Stronger than covalent and metallic bonds.	Weaker than ionic and stronger than metallic bonds.	Weaker than ionic and covalent bonds.
Electronegativity	Influences the strength of the bond.	Influences the strength of the bond.	Little or no influence on the strength of the bond.
Character of Bond	Non-directional	Directional	Non-directional
Existence in the Physical State	Solid	Solid, liquid, and gas	Solid except mercury
Electrical Conductivity	Have a low conductivity	Have a very low conductivity	Have a very high conductivity
Ductility	Not ductile	Not ductile	Ductile
Melting Point	Have higher melting points	Have lower melting points	Have high melting points
Boiling Point	Have higher boiling points	Have lower boiling points	Have high boiling points
Examples	NaCl, CaO, KBr	H_2, N_2, H_2O, CH_4	Fe, Au, Ag

4.5 Chemical Bonding and Compound Properties

4.5.1 Comparison of Ionic Bonds and Metallic Bonds

The comparison of ionic bonds and metallic bonds involves understanding the different ways in which atoms interact to form compounds.

1. Ionic Bonds

Formation: Ionic bonds are formed by the transfer of electrons between atoms, which leads to the creation of positive and negative ions.

Strength: Ionic compounds have high melting and boiling points due to the strong attraction between the positively and negatively charged ions within the compound.

Conductivity: Ionic compounds do not conduct electricity in a solid state because the ions are fixed in a rigid structure. However, when the compound is melted or dissolved in water, the ions become mobile and can conduct electricity.

Uses: Ionic compounds such as sodium chloride ($NaCl$) are used in cooking, whereas calcium carbonate ($CaCO_3$) is employed in construction materials.

2. Metallic Bonds

Formation: Metallic bonds consist of a "sea" of delocalized electrons shared among metal atoms.

Strength: Metallic bonds have considerable strength, which enhances the significant tensile strength and hardness of metals.

Conductivity: Metals are excellent conductors of heat and electricity due to their delocalized electrons.

Uses: Copper is used in electrical wiring, while steel is utilised in construction because of its strength and durability.

AMAZING FACT

Gold can be hammered into sheets thinner than human hair because of the free movement of delocalised electrons in metallic bonds.

4.5.2 Melting and Boiling Points of Ionic and Covalent Compounds

1. Ionic Compounds

Melting Point: Ionic compounds have high melting points because of their strong ionic bonds. A substantial amount of energy is needed to break these bonds and change from a solid to a liquid state.

Boiling Point: Ionic compounds also possess high boiling points due to the strong attractive forces between ions in the liquid state. A significant amount of energy is required to overcome these forces and transition to the gaseous state.

2. Covalent Compounds

Melting Point: Covalent compounds generally have lower melting points than ionic compounds. This is due to the weaker forces between covalent molecules, which lead to lower energy requirements for melting.

Boiling Point: Covalent compounds also show lower boiling points than ionic compounds. The weaker forces facilitate an easier transition from the liquid to a gaseous state.

4.6 Conductivity of Ionic Compounds, Covalent Compounds, and Metallic Bonds

1. Ionic Compounds

In the solid state, ionic compounds do not conduct electricity because the ions remain fixed in the lattice structure. However, in the molten or dissolved state, they can conduct electricity as the ions can move freely.

2. Covalent Compounds

Covalent compounds are poor conductors of electricity due to their lack of free charged particles and their limited ability to conduct electricity in any state.

3. Metallic Bonds

Metallic bonds are excellent conductors of electricity because the delocalised electrons can move freely within the metallic lattice, allowing for efficient conduction of electricity.

MISCONCEPTION Vs REALITY

Myth: Covalent compounds do not conduct electricity.

Reality: While most covalent compounds are poor conductors, some, like graphite, conduct electricity due to delocalised electrons.

4.6.1 Properties in Terms of Bonding and Structure

Melting and boiling points are influenced by the strength and type of intermolecular forces. Ionic compounds have high melting and boiling points due to strong ionic bonds, whereas covalent compounds typically show lower melting and boiling points. Solubility depends on the nature of bonding and molecular polarity, with polar and ionic compounds frequently dissolving in polar solvents such as water. Conductivity is determined by free-charged particles, with ionic compounds conducting electricity when ions are mobile, while covalent compounds are generally poor conductors.

Keep in Mind

Substances that ionise in water and conduct electricity are known as electrolytes. For example, when hydrochloric acid is dissolved in water, it ionises into hydrogen ions and chloride ions, which can conduct electricity.

$$HCl \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

Other examples of electrolytes are strong acids, strong bases, and salts.

Ionic compounds are hard and have high strength due to their strong ionic bonds, whereas the hardness and strength of covalent compounds vary depending on the bond types and atomic arrangements. Metallic compounds show malleability and ductility. Reactivity is affected by bond types and atomic arrangements, with ionic compounds typically being more reactive than metallic compounds, which tend to be less reactive.

4.7 Industrial Uses of Graphite, Diamond and Metals

4.7.1 Graphite

i) Graphite is a material capable of conducting electricity and is frequently used as electrodes for batteries and electrical wires.

ii) It also has good thermal conductivity, making it valuable for thermal sinks, crucibles, and refractories.

iii) It is also used in the aerospace industry due to its high thermal stability.

iv) Graphite has lubricating properties, making it beneficial in lubricants, decreasing friction and wear in items such as locks, pencils, and bearings.

Real-Life Application

Graphite is used in rechargeable lithium-ion batteries for smartphones and laptops because it facilitates the easy movement of ions in and out.

4.7.2 Diamond

i) Diamond is the most rigid material, making it ideal for cutting, grinding, and drilling tools in mining, construction, and machining industries.

ii) It also shows high thermal conductivity and insulation capabilities, making it ideal for use in heat sinks and the semiconductor industry, where high-performance electronic components are manufactured.

iii) Diamond's high refractive index makes it useful for optical applications such as lenses, prisms, and cutting tools.

iv) It is resistant to most chemicals, which makes it suitable for use in demanding chemical environments.

FUN FACT

Diamonds naturally form only under extreme pressure and temperature, approximately 150–200 km beneath the Earth's surface. Scientists can now produce synthetic diamonds in laboratories within days using advanced high-pressure and high-temperature techniques.

4.7.3 Metals

i) Metals are strong and rigid, making them suitable for construction and manufacturing.

ii) They have widely electrical conductivity, making them suitable for electrical conductors, wires, and cables.

iii) They are malleable and ductile, enabling them to be shaped into various forms. They are widely used in manufacturing components across various industries such as automotive, aerospace, and consumer electronics.

iv) They also have high thermal conductivity, making them suitable for applications such as heat exchangers and electronic components.

4.8 Dot and Cross Diagrams

Dot and Cross Diagrams use dots and crosses to represent the arrangement of outer shell electrons in an ionic or covalent compound. All electrons are represented by dots (•) and crosses (×). They show the total number of lone pairs present in each atom and are useful for understanding electron distribution in compounds.

4.8.1 Steps for Drawing the Dot and Cross Diagrams

i) Write the symbols for each atom or ion.

ii) Identify the central atom and arrange the other atoms around it. The central atom is defined as the one with the highest number of bonds.

iii) Write the electronic configuration for each atom or ion, using dots and crosses.

iv) Arrange eight electron positions around each atom.

v) Illustrate the transfer or sharing of electrons using arrows.

vi) Indicate the charge of each ion using brackets. The charge for each ion is placed at the upper right corner.

vii) Ensure that each atom, apart from hydrogen, is paired with eight electrons, whether in lone or shared pairs.

4.9 Lewis Dot Structure

Lewis dot structures, on the other hand, represent the distribution of outer shell electrons around an atom or molecule. They predict electron pairs and bonds within compounds, using dots to represent electrons and lines to represent bonds.

4.9.1 Steps for Drawing the Lewis Dot Diagrams

i) Draw a skeletal diagram linking all atoms with single bonds.

ii) Assign lone pairs of electrons to each atom.

iii) If needed, draw double or triple bonds to satisfy octet configurations.

iv) Surround each atom with eight electrons (except for hydrogen).

v) Represent ionic bonds as full valence shells, where the more electronegative element gains electrons.

vi) Use brackets to enclose charged atoms and superscript the charge, if necessary.

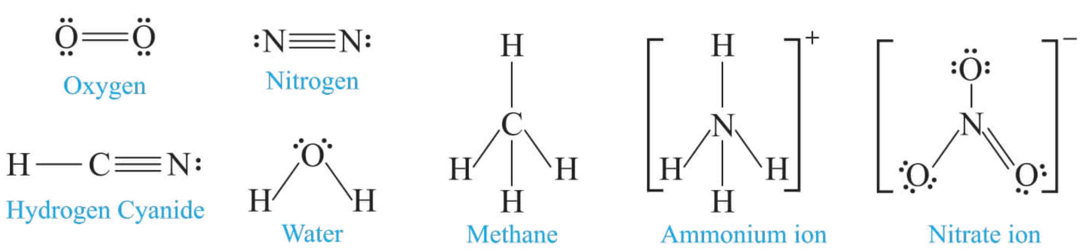


Fig. 4.2: Lewis dot structures for molecules and ions

ACITIVITY

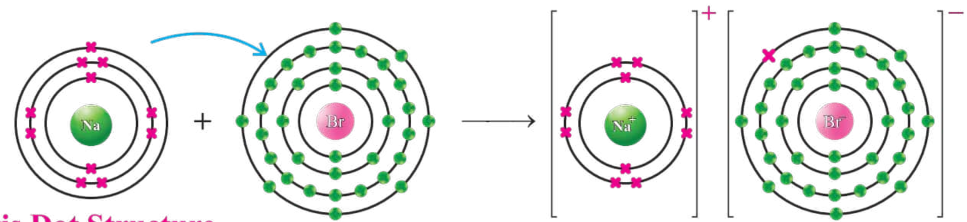
Use coloured beads to represent ionic and covalent compounds. Compare the rigid three-dimensional lattice of NaCl with the simple molecular structure of CH_4.

4.9.2 Dot and Cross Models of Ionic Compounds

Ionic bonds in binary compounds such as NaBr, NaF, and $CaCl_2$ using dot-and-cross diagrams and Lewis dot structures.

1. Sodium Bromide (NaBr)

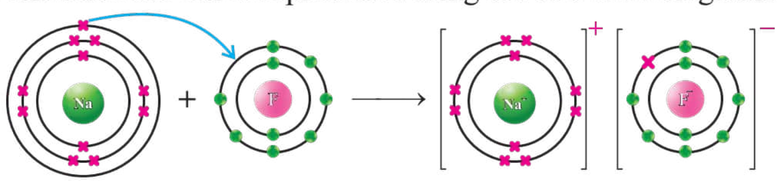
Sodium (Na) has one electron in its outer shell, which is transferred to bromine (Br). The resultant sodium ion (Na^+) and bromide ion (Br^-) are held together by ionic bonds. This structure can be represented using dot-and-cross diagrams.



Lewis Dot Structure

2. Sodium Fluoride (NaF)

Sodium (Na) possesses one electron in its outer shell, which is transferred to a fluorine (F) atom. The resulting sodium ion (Na^+) and fluoride ion (F^-) are held together by ionic bonds. This structure can be represented using dot-and-cross diagrams.



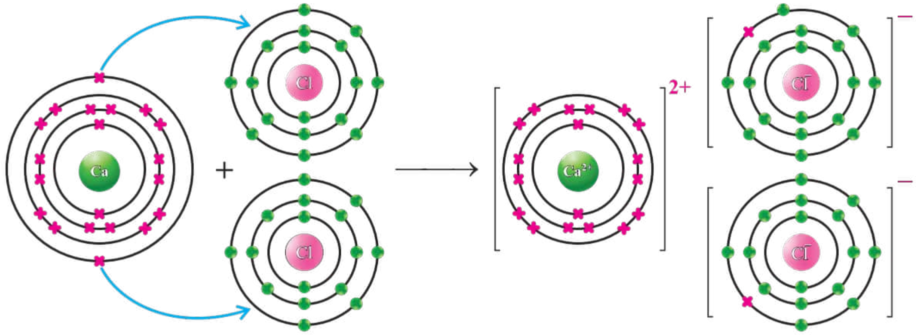
Lewis Dot Structure

3. Calcium Chloride ($CaCl_2$)

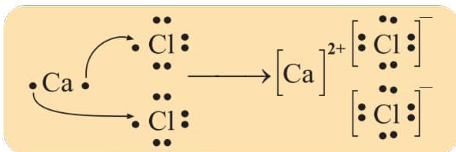
Calcium (Ca) has two electrons in its outer shell, which are transferred to chlorine (Cl). The resulting calcium ion (Ca^{2+}) and chloride ion (Cl^-) are held together by ionic bonds. The structure can be represented using dot-and-cross diagrams.

CHEMISTRY IN LIFE

The salt you sprinkle on food and the calcium chloride used to melt ice on roads are both examples of ionic compounds.



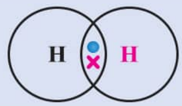
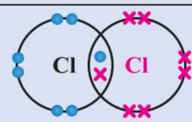
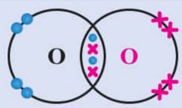
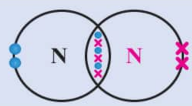
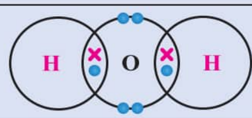
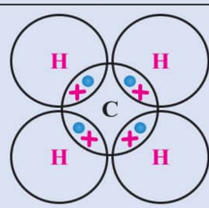
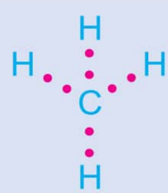
Lewis Dot Structure

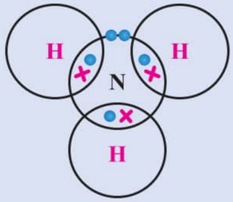
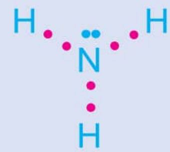
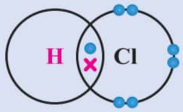
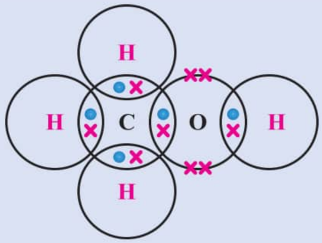
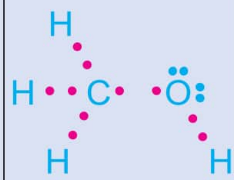
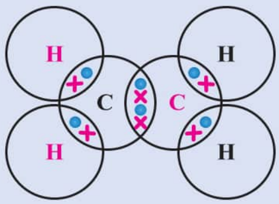
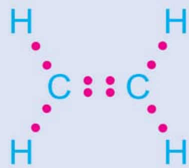
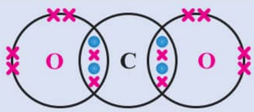
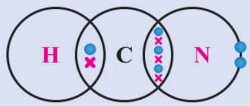


4.9.3 Dot and Cross Models of Covalent Compounds

Covalent bonds in simple molecules such as H_2, O_2, N_2, H_2O, NH_3, HCl, CH_3OH, C_2H_4, CO_2, and HCN can be illustrated using dot-and-cross diagrams and Lewis-dot structures.

Table 4.4: Dot-and-cross diagrams and lewis structures of simple molecules

Name	Formula	Dot and Cross Diagram	Lewis Structure
Hydrogen	H_2		
Chlorine	Cl_2		
Oxygen	O_2		
Nitrogen	N_2		
Water	H_2O		
Methane	CH_4		

Ammonia	NH_3		
Hydrogen chloride	HCl		
Methyl alcohol	CH_3OH		
Ethene	C_2H_4		
Carbon dioxide	CO_2		
Hydrogen cyanide	HCN		

KEY POINTS

1. A chemical bond is the force that holds atoms together to form compounds.
2. Ionic bonds occur through electron transfer, resulting in oppositely charged ions that attract one another.
3. Covalent bonds are formed by sharing electrons between atoms to fill their outer shells.
4. Atoms follow the octet or duplet rule to attain stable noble gas configurations.
5. Noble gases in Group 18 have complete outer electron shells, making them stable and unreactive.
6. Metallic bonds involve delocalised electrons moving freely within a lattice of positive metal ions, resulting in conductivity, malleability, and ductility in metals.

7. A coordinate covalent bond (dative bond) occurs when both shared electrons come from a single atom, for example, NH_4^+, H_3O^+.
8. Lewis dot structures use dots to represent valence electrons and visualise bonding.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The primary reason atoms form chemical bonds is to:
   (a) become unstable
   (b) increase their atomic number
   (c) achieve a stable electronic configuration
   (d) become radioactive

ii) The rule that indicates atoms generally have eight electrons in their outermost shell is:
   (a) octet rule (b) duplet rule
   (c) triplet rule (d) hexlet rule

iii) The group of elements least likely to form chemical bonds because of their filled valence shells are:
   (a) halogens (b) noble gases
   (c) alkali metals (d) transition metals

iv) The charge on a cation is:
   (a) positive (b) negative
   (c) neutral (d) variable

v) The type of bond formed by the transfer of electrons between atoms with a significant difference in electronegativity is:
   (a) covalent bond (b) metallic bond
   (c) ionic bond (d) coordinate covalent Bond

vi) The state in which ionic compounds typically exist at room temperature is:
   (a) gas (b) liquid
   (c) solid (d) plasma

vii) The property of metals that makes them good conductors of electricity is:
   (a) high electronegativity (b) low melting point
   (c) delocalized electrons (d) covalent bonding

viii) The type of bond formed by the equal or mutual sharing of electrons between two atoms is:
(a) ionic bond (b) metallic bond
(c) covalent bond (d) hydrogen bond

ix) The primary industrial use of graphite is:
(a) construction material
(b) electrical conductor in electrodes
(c) fertiliser additive
(d) food preservative

x) Metals are widely used in electrical wiring due to their:
(a) high melting point (b) electrical conductivity
(c) shiny appearance (d) high density

B. Restricted Response Questions (RRQs)

i) Why do representative elements in groups VA-VIIA tend to gain electrons?
ii) Provide examples of substances that can ionise when dissolved in water.
iii) Explain why the bond in an oxygen molecule is called a double bond.
iv) Differentiate between single, double, and triple covalent bonds.
v) Why do noble gases show low reactivity?.
vi) Solid sodium chloride does not conduct electricity, whereas a solution of sodium chloride does. why?

C. Extended Response Questions (ERQs)

i) What are the octet and duplet rules, and how do they relate to the stability of noble gas electronic configurations?
ii) How are ions formed? Explain the difference between cations and anions, providing examples.
iii) Compare ionic, covalent, and metallic bonds with respect to their formation and key properties.
iv) Why do ionic compounds possess high melting and boiling points? What factors influence these properties?
v) What are the industrial applications of graphite, diamond, and metals? Explain how their properties are suited to these applications.

UNIT 5

STOICHIOMETRY

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-42 State the formulae of compounds.

B-43 Define molecular formula of a compound as the number and type of different atoms in one molecule.

B-44 Define empirical formula of a compound as the simplest whole number ratio of different atoms in a molecule.

B-45 Deduce the formula and name of binary ionic compounds from ions given relevant information.

B-46 Deduce the formula of a molecular substance from the given structure of molecule.

B-47 Define mole as amount of substance containing Avogadro's number ($6.02x10^{23}$) of particles.

B-48 Explain the relationship between a mole and Avogadro's number.

B-49 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.

B-50 Construct chemical equations and ionic equations to show reactants forming products, including state symbols.

B-51 Deduce the symbol equation with state symbols for a chemical reaction given relevant information.

Introduction

Stoichiometry is a branch of chemistry that focuses on the relationships between the quantities of substances involved in a chemical reaction. The term "stoichiometry" is derived from two Greek words: "stoicheion", meaning element, and "metron", meaning measure. It highlights the importance of quantitative measurements of both reactants and products in chemical processes.

Stoichiometry involves the study of balanced chemical equations to determine the quantities of reactants and products involved in a chemical reaction. This knowledge is essential for calculating the amounts of products formed in a reaction, including their masses, moles, and volumes.

CHEMISTRY IN LIFE

Stoichiometry is used in medicine to determine appropriate drug dosages. A small miscalculation can lead to overdose or ineffectiveness.

5.1 Chemical Formula

The symbolic representation of a molecule or compound is known as a chemical formula. It uses atomic symbols and numerical subscripts to indicate the number and types of atoms present in a substance. For example, the chemical formula of sodium chloride is NaCl, which represents one atom of sodium and one atom of chlorine. In contrast, the chemical formula of nitric acid is HNO_3, denoting one atom of hydrogen, one atom of nitrogen, and three atoms of oxygen. There are two main types of chemical formulae molecular and empirical formulae.

Keep in Mind:

Chemical formulae use symbols from the periodic table to represent the elements present in the formula. The number of each type of atom present in the formula is indicated by subscripts placed next to the element symbols.

5.1.1 Molecular Formula

The molecular formula indicates the exact number of atoms in a molecule. It specifies the number of atoms of each element present in a given molecule, using subscripts to represent the quantities. For example, the chemical formula for water is H_2O. In this formula, the subscript two (2) indicates that there are two hydrogen atoms, while the absence of a subscript for oxygen indicates a single oxygen atom. Another illustration is carbon dioxide (CO_2), where the lack of a subscript for carbon means one carbon atom, and the subscript two (2) for oxygen indicates two oxygen atoms.

5.1.2 Empirical Formula

The formula that shows the simplest whole-number ratio of different atoms in a compound is called the empirical formula. This formula represents the relative number of atoms of each element in the compound. For example, the molecular formula of glucose is $C_6H_{12}O_6$, indicating six carbon atoms, twelve hydrogen atoms, and six oxygen atoms. The simplest whole-number ratio between carbon, hydrogen, and oxygen is 1:2:1; hence, the empirical formula of glucose is CH_2O.

Do You Know?

The word empirical derives from the Greek word empeirikos, meaning experimental. The empirical formula is named as such because it is obtained from experiments that determine the proportion of each element in a compound.

Table 5.1: Compounds with molecular and empirical formulae

Name of Compound	Molecular Formula	Nature of Compound	Empirical Formula/ Formula Unit
Water	H_2O	Covalent	H_2O
Ammonia	NH_3	Covalent	NH_3
Methane	CH_4	Covalent	CH_4
Nitric acid	HNO_3	Covalent	HNO_3
Acetic acid	CH_3COOH	Covalent	CH_2O
Glucose	$C_6H_{12}O_6$	Covalent	CH_2O
Acetylene	C_2H_2	Covalent	CH
Benzene	C_6H_6	Covalent	CH
Sulphuric acid	H_2SO_4	Covalent	H_2SO_4
Sodium chloride	NaCl	Ionic	NaCl
Sodium hydroxide	NaOH	Ionic	NaOH
Calcium hydroxide	$Ca(OH)_2$	Ionic	$Ca(OH)_2$
Copper sulphate	$CuSO_4$	Ionic	$CuSO_4$
Sodium carbonate	Na_2CO_3	Ionic	Na_2CO_3

5.2 Determination of Formula and Name of Binary Ionic Compounds from Ions

Compounds made of atoms from two different elements are known as binary compounds. For example, sodium chloride is a binary compound composed of atoms of two different elements: sodium and chlorine. Similarly, ammonia contains atoms of nitrogen and hydrogen, making it a binary compound. To determine the formula and name of binary ionic compounds from their ions, it is essential to know the charges of cations and anions.

5.2.1 Steps to Deduce the Formula and Name of binary ionic compounds

i) To deduce the formula and name of binary ionic compounds, you first need to identify the cations (typically metals) and anions (generally nonmetals) present in the compounds.

ii) Determine the charges of cations and anions, which are often related to the group number of the elements. For main group elements, the charge typically corresponds to the group number. For example, elements in group IA have a +1 charge, those in group IIA have a +2 charge, elements in group VIA have a –2 charge, and elements in group VIIA have a –1 charge, and so forth. The charges of the cation and anion become the subscripts of one another in the formula of a compound, and these subscripts should be simplified to the lowest possible ratio.

iii) The name of the cation comes before the name of the anion. The name of the cation remains the same as the name of the element, whereas the name of the anion is modified by replacing the ending of the element name with "-ide." Sometimes it is "-ite" or "-ate."

Examples

i) A compound consists of sodium ions (Na) and fluoride ions (F). The formula of the compound is NaF, and the name of the compound is sodium fluoride.

ii) A compound that contains iron(III) ions (Fe^{3+}) and oxide ions (O^{2-}) has the chemical formula Fe_2O_3. This compound is known as iron(III) oxide or ferric oxide. The "III" in iron(III) oxide indicates that iron has a +3 oxidation state.

Table 5.2: Some examples of binary ionic compounds, along with the relevant ions, formulae, and names

Cation	Anion	Charge on Cation	Charge on Anion	Formula	Name of Compound
Sodium (Na^+)	Chloride (Cl^-)	+1	−1	NaCl	Sodium chloride
Magnesium (Mg^{2+})	Oxide (O^{2-})	+2	−2	MgO	Magnesium oxide
Calcium (Ca^{2+})	Fluoride (F^-)	+2	−1	CaF_2	Calcium fluoride
Aluminium (Al^{3+})	Bromide (Br^-)	+3	−1	$AlBr_3$	Aluminum bromide
Magnesium (Mg^{2+})	Nitride (N^{3-})	+2	−3	Mg_3N_2	Magnesium nitride
Iron(II) (Fe^{2+})	Sulphide (S^{2-})	+2	−2	FeS	Iron(II) sulphide

Self-Assessment

i) How does the empirical formula differ from a molecular formula?
ii) Explain why glucose ($C_6H_{12}O_6$) and acetic acid (CH_3COOH) have the same empirical formula.
iii) If a compound has the empirical formula CH, what types of compounds could it represent?

5.3 Deducing Molecular Formulas from Molecular Structures

To deduce the formula of a molecular substance from the structure of its molecules, a systematic approach involves the following steps:

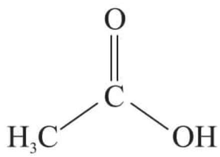
i) **Identify the Atoms Present:** Examine the molecular structure and identify the symbols that represent different elements. Take note of each type of atom present within the molecule.

ii) **Count the Number of Each Type of Atom:** Analyse the molecular structure to determine the quantity of each type of atom. Pay attention to the subscripts and coefficients associated with each atom in the structure, as they indicate the number of atoms present. For example, consider the molecule represented by the structural formula:

iii) **Identify Elements:** The structural formula has carbon (C), hydrogen (H), and oxygen (O).

iv) **Count the Atoms:** The structural formula contains 2 atoms of C, 4 atoms of H, and 2 atoms of O.

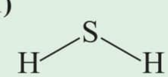
So, from the structural formula, the deduced molecular formula is $C_2H_4O_2$.



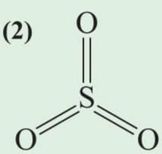
Self-Assessment

Consider various structural formulae of sulphur. Outline the steps you would take to deduce the molecular formula from this structure, highlighting the importance of counting and simplifying atom ratios.

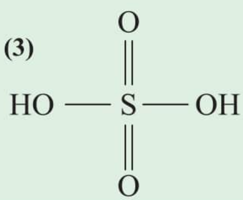
(1)



(2)



(3)



ACITIVITY

Exploring Structural Formulas: Building and Drawing Molecules

Objectives: To familiarize students with molecular formulas, utilize hands-on modeling kits for constructing diverse molecular structures, reinforce understanding through drawing, and promote independent exploration into various compound structural formulas.

Materials: Writing board, instruction sheet, modeling kits or beads and toothpicks, and worksheets with different structural formulas (e.g., MgO, CH_4, C_2H_6, C_3H_6, C_6H_6)

Procedure: Students visually interpret molecular formulas from a chart, construct models using kits, share and discuss structures, draw on the board, learn complex arrangements, and explore additional compounds for homework, fostering research and understanding.

Self-Assessment

i) What is the main difference between a molecular formula and an empirical formula?

ii) If the empirical formula of a compound is CH_2 and its molecular mass is 28 amu, what is its molecular formula?

5.4 Mole and Avogadro's Number

We know that atoms are extremely small particles, making it impossible to weigh or count them. That is why we increase the sizes of samples to the point where we can weigh or count them. Therefore, they cannot be counted in pairs, dozens, or reams. The chemist's unit of measurement is called a mole.

Amedeo Avogadro (Italian Scientist) 1776 — 1856

5.4.1 Mole

The word mole means a 'huge mass'. Its symbol is 'mol' and is represented by 'n'. One mole is the amount of substance that contains as many particles (atoms, molecules, ions, or formula units) as the number of atoms in exactly 12 g of carbon-12.

5.4.2 Avogadro's Number

One mole contains 6.0221421×10^{23} particles, which we usually round to 6.022×10^{23}. Scientists call this value Avogadro's number in honour of the Italian scientist Amedeo Avogadro (1776-1856). The unit of Avogadro's number is read as "particles per mole", and is represented by N_A. The particles can be atoms, molecules, ions, or electrons.

Interesting Information

The concept of the mole was first proposed by Wilhelm Ostwald in 1895, while Amedeo Avogadro's law was later developed and refined by Johann Josef Loschmidt and Josef Perri in 1909.

5.4.3 Relationship Between Mole and Avogadro's Number

The relationship between a mole and Avogadro's number is very important in chemistry. One mole of any substance contains 6.022×10^{23} tiny particles. Avogadro's number helps us connect the very small world of atoms and molecules to the amounts we can measure in the lab. For example, one mole of water has 6.022×10^{23} water molecules, and one mole of sodium has the same number of sodium atoms.

Table 5.3: Names, formulae, molar masses, and number of particles of various substances

Name of Substance	Formula	Mass (amu)	Mass of one mole (g/mol)	Number of particles in one mole
Oxygen atom	O	16	16	6.022×10^{23} atoms
Oxygen molecule	O_2	32	32	6.022×10^{23} molecules
Water	H_2O	18	18	6.022×10^{23} molecules
Potassium nitrate	KNO_3	101	101	6.022×10^{23} formula units
Carbonate ion	CO_3^{-2}	60	60	6.022×10^{23} ions

Interesting Information

If all 6 billion people on Earth were to do nothing but count the gumballs in one mole at a rate of one gumball per second, it would take over 3 million years to count all the gumballs!

5.5 Molar Mass

The mass of one mole of a substance (atoms, molecules, or formula units), expressed in grams, is known as its molar mass. The unit for molar mass is grams per mole (g/mol). To determine the molar mass, we convert the units from atomic mass units to grams of the substance. The substance may be an element or a compound. For example, a hydrogen atom has an atomic mass of 1.008 amu, so the molar mass of a hydrogen atom is 1.008 g/mol, and it contains 6.022×10^{23} atoms of hydrogen. On the other hand, ammonia (NH_3) has a molecular mass of 17 amu, so the molar mass of ammonia is 17 g/mol, and it contains 6.022×10^{23} molecules of ammonia.

Table 5.4: Molar masses of elements and compounds

Name	Symbol / Formula	Nature		Atomic/ Molecular/ Formula Mass (amu)	Molar Mass (g/mol)	Number of Particles /mol
Carbon	**C**	Element		12.0	12.0	6.022×10^{23} atoms
Sulphur	**S**	Element		32.0	32.0	6.022×10^{23} atoms
Mercury	**Hg**	Element		200.6	200.6	6.022×10^{23} atoms
Carbon dioxide (dry ice)	**CO_2**	Compound		44.0	44.0	6.022×10^{23} molecules
Sodium chloride	**NaCl**	Compound		58.5	58.5	6.022×10^{23} formula units
Copper(II) sulphate	**$CuSO_4$**	Compound		159.6	159.6	6.022×10^{23} formula units

Example 5.1 **Calculate the molar mass of hydrogen peroxide (H_2O_2).**

SOLUTION

The molar mass of hydrogen peroxide can be calculated as:

2 moles of hydrogen	×	1g/mol	= 2g
2 moles of oxygen	×	16g/mol	= 32g
The molar mass of hydorgen peroxide, H_2O_2			= 34g

Hence, the molar mass of hydrogen peroxide is 34g/mol

Test Yourself: Calculate the molar masses of benzene (C_6H_6) and glucose ($C_6H_{12}O_6$).

5.6 Calculations Using the Relationship Between Amount of Substance, Mass, and Molar Mass

5.6.1 Calculation the Number of Moles of a Substance

If we know the mass of a substance, we can calculate the number of moles by dividing the mass by the molar mass.

$$\text{Moles of a Substance} = \frac{\text{Given mass of a Substance (g)}}{\text{Molar Mass (g/mol)}}$$

Example 5.2 **How many moles are there in 27g of H_2O?**

SOLUTION

Mass of H_2O = 27g

Moles of H_2O = Unknown

Molar Mass of H_2O = 18g/mol

$$\text{Moles of } H_2O = \frac{\text{Mass of } H_2O}{\text{Molar Mass of } H_2O} = \frac{27\text{g}}{18\text{g/mol}} = 1.5\text{mol}$$

Test Yourself: Calculate the number of moles in 19.6g of H_2SO_4 and 4.6g of sodium (Na).

5.6.2 Calculation of Mass of a Substance

You can determine the mass of a specified number of moles of a substance by rearranging the mole equation:

$$\text{Mass of Substance} = \text{Moles of a Substance} \times \text{Molar Mass}$$

Example 5.3 **Calculate the mass, in grams, of 10moles of H_2O.**

SOLUTION

Molar mass of H_2O = 2(1) + 1(16) = 2 + 16 = 18g/mol

Mass of H_2O = Molar mass of H_2O × Moles of H_2O = 18g/mol × 10mol = 180g

Test Yourself: Calculate the mass of 2.50moles of glucose.

5.6.3 Calculation of Molar Mass

If we know the mass and the number of moles of a substance, we can calculate the molar mass by dividing the mass by the number of moles.

$$\text{Molar Mass of a Substance} = \frac{\text{Given Mass(g)}}{\text{Moles of a Substance}}$$

Example 5.4 **What is the molar mass of carbon dioxide if 22 grams of it equals 0.5 moles?**

SOLUTION

Mass of carbon dioxide (CO_2) = 22g

Moles of carbon dioxide = 0.5mol

Molar mass of carbon dioxide = ?

Molar mass of carbon dioxide $= \frac{22g}{0.5mol} = 44g/mol$

Test Yourself: A sample of copper has a mass of 31.8grams and contains 0.5 moles. What is its molar mass?

5.6.4 Calculation of the Number of Particles in a Substance

If we know the number of moles of a substance, we can calculate the number of particles (atoms, molecules, or formula units) by multiplying the number of moles by Avogadro's number.

$$\text{Number of Particles} = n \times N_A$$

Example 5.5 **If a sample contains two moles of carbon dioxide, how many molecules does it have?**

SOLUTION

Moles of carbon dioxide (CO_2) = 2 mol

Molecules of carbon dioxide = ?

Avogadro's number (N_A) = 6.02×10^{23} molecules/mol

Molecules of carbon dioxide = $2\text{ mol} \times 6.02 \times 10^{23}$ molecules/mol

= 12.04×10^{23} molecules

Test Yourself: Suppose you have 3moles of fructose. How many molecules are present in this amount?

Chemistry in Your Body

Stoichiometry shows how much oxygen your body needs. For example, complete combustion of glucose ($C_6H_{12}O_6$) in cells requires exactly 6 moles of O_2 for each mole of glucose.

5.7 Construction of Chemical Equations

A chemical equation is the symbolic representation of a chemical reaction. In a chemical equation, the starting substances, or reactants, are on the left; the final substances, or products, are on the right, with an arrow placed between them to indicate a transformation. The numbers and types of atoms are the same on both sides of the reaction arrow, as required by the law of conservation of mass. The chemical equation in which the number of atoms of each element is equal on both sides is called a balanced chemical equation. The symbols in subscript represent the physical state of the substance.

Table 5.5: Symbols commonly used in chemical equations

Symbol	Meaning
+	Plus or added to (placed between substances)
$\longrightarrow$	Yields; produces (points to products)
(*s*)	Solid state (written after a substance)
(*l*)	Liquid state (written after a substance)
(*g*)	Gaseous state (written after a substance)
(*aq*)	Aqueous solution (substance dissolved in water)
Δ	Heat is added (when written above or below arrow)

Examples of Balanced Chemical Equations

1. Combustion of Methane

Consider the molecular reaction between one mole of methane (CH_4) and two moles of oxygen (O_2) to produce one mole of carbon dioxide (CO_2) and two moles of water vapour (H_2O).

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

2. Decomposition of Calcium Carbonate

Calcium carbonate ($CaCO_3$) decomposes upon heating to form calcium oxide (CaO) and carbon dioxide(CO_2).

$$CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$$

A balanced chemical equation shows the reactants and products of a reaction, including their correct chemical formulas and the corresponding mole ratios. It displays the conservation of mass by maintaining equal quantities of each type of atom on both sides. Moreover, when state symbols are included, it indicates the physical states of the substances involved.

Self-Assessment

Write the chemical equation for the decomposition of calcium carbonate, including state symbols.

Test Yourself:

Write the molecular equation for the decomposition of Sodium carbonate, including state symbols.

3. Neutralisation Reaction Between Hydrochloric Acid and Sodium Hydroxide

Hydrochloric acid reacts with sodium hydroxide to produce sodium chloride and water:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

The full ionic equation for the reaction is:

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

The sodium (Na^+) and chloride (Cl^-) ions are spectator ions (ions that do not participate in the reaction and appear unchanged on both sides). Therefore, the net equation is:

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

KEY POINTS

1. Stoichiometry is the branch of chemistry that deals with the quantitative relationships between reactants and products in a chemical reaction.
2. A chemical formula is the symbolic representation of the composition of a compound.
3. A molecular formula specifies the exact number of each type of atom in a molecule, such as $C_6H_{12}O_6$ for glucose.
4. An empirical formula shows the simplest whole-number ratio of atoms in a compound, such as CH_2O for glucose.
5. The molar mass is the mass of one mole of a substance expressed in grams per mole; for example, the molar mass of carbon is 12g/mol, and that of carbon monoxide (CO) is 28 g/mol.
6. Avogadro's number is 6.022×10^{23} and represents the number of atoms, ions, or molecules in one mole of a substance; it is named after the Italian scientist Amedeo Avogadro.
7. The formula mass of a compound is the sum of the atomic masses of all atoms present in its formula unit.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) Stoichiometry deals with:
   - (a) colour changes
   - (b) physical properties
   - (c) quantitative relationships
   - (d) speed of reaction

ii) The molecular formula of glucose is:
   - (a) C_6H_6
   - (b) $C_6H_{12}O_6$
   - (c) C_6H_{12}
   - (d) CO_2

iii) The type of formula that shows the simplest whole-number ratio of different atoms in a compound is:
   - (a) molecular formula
   - (b) empirical formula
   - (c) chemical formula
   - (d) structural formula

iv) The number of particles present in ten mole of any substance is:
   - (a) 6.022×10^{23}
   - (b) 6.022×10^{23}
   - (c) 6.022×10^{24}
   - (d) 6.022×10^{25}

v) The symbol that represents a solid state in a chemical equation is:

(a) (s) (b) (ℓ)
(c) (g) (d) (aq)

vi) If a sample of copper has a mass of 31.8 grams and contains 0.5 moles, its molar mass is:

(a) 15.9 g/mol (a) 31.8 g/mol
(c) 63.6 g/mol (d) 0.0159 g/mol

vii) In the formula CaF_2, the charge on the fluoride ions is:

(a) +1 (b) −1
(c) +2 (d) −2

viii) The molar mass of water (H_2O) is:

(a) 18 g/mol (b) 22.414 dm^3/mol
(c) 6.022×10^{23} g/mol (d) 2 g/mol

ix) The term that represents the actual number of atoms in a molecule is:

(a) Empirical formula (b) Molecular formula
(c) Chemical formula (d) Structural formula

x) The chemical formula for a compound composed of magnesium ions (Mg^{2+}) and chloride ions (Cl^-) is:

(a) $MgCl_2$ (b) MgCl
(c) Mg_2Cl_2 (d) Mg_2Cl

B. Restricted Response Questions (RRQs)

i) Define the term Stoichiometry.

ii) What is the chemical formula for butane, and how many hydrogen atoms does it contain?

iii) Give the empirical formula of glucose ($C_6H_{12}O_6$).

iv) State the difference between the terms 'mole' and 'molar mass'.

v) Calculate the molar mass of hydrogen sulphide (H_2S).

vi) How many molecules are present in 1.5 moles of ammonia?

vii) What information does a balanced chemical equation provide?

C. Extended Response Questions (ERQs)

i) Discuss the significance of stoichiometry in chemical reactions, providing examples.

ii) Define the empirical formula of a compound with examples and explain how to determine it from a given molecular formula.

iii) Define a mole. Explain its significance in chemistry, and relate it to Avogadro's number.

D. Numerical Problems

i) Calculate the number of moles in 39.2g of H_2SO_4 (molar mass = 98g/mol).

ii) A sample of copper has a mass of 15.9g and contains 0.25moles. What is its molar mass?

iii) Calculate the relative molecular mass of sulphuric acid (H_2SO_4).

iv) Determine the relative formula mass of aluminium sulphate ($Al_2(SO_4)_3$).

v) If you have 2.5moles of sodium bromide, determine the number of formula units in this sample.

UNIT
6

ELECTROCHEMISTRY

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-52 Define redox reactions as simultaneous oxidation and reduction in terms of oxygen, hydrogen, electrons and changes in oxidation state.

B-53 Use Roman numerals to indicate oxidation number of an element in a compound.

B-54 Identify oxidizing and reducing agents in a redox reaction in term of electron (s).

B-55 Recognize that the oxidation number of elements in their free state is zero.

B-56 Derive the formula of ionic compounds from ionic charges and oxidation numbers.

B-57 Identify that the oxidation number of a monoatomic ion is the same as the charge on the ion.

B-58 Explain that the sum of the oxidation numbers in a neutral compound is zero .

B-59 Explain that the sum of the oxidation numbers in an ion is equal to the charge on the ion

B-60 Identify redox reactions by the colour changes involved when using acidified aqueous Potassium manganate (VII) to (II) and aqueous potassium iodide .

B-61 Define corrosion and discuss methods to prevent it.
(some examples may include barrier method such as using paint, galvanizing, electroplating; sacrificial protection such as using magnesium blocks in ships.)

Introduction

Electrochemistry is a branch of chemistry that studies the conversion of chemical energy into electrical energy, as well as the chemical changes caused by an electric current. It primarily involves oxidation–reduction (redox) reactions, which can either occur spontaneously to produce electricity or require an external electric supply to proceed. It explains practical phenomena such as corrosion and its prevention, along with applications like batteries, electroplating, and metal refining.

Do You Know?

Alessandro Volta developed the first battery in 1800 using zinc and copper plates separated by a saline cloth. In 1836, John Frederic Daniell designed a more reliable version of the Daniell cell, which powered early telegraph systems and revolutionised global communication.

6.1 Redox Reactions

In a redox reaction, both oxidation and reduction processes occur simultaneously. (Oxidation involves the loss of electrons by one substance, while reduction involves the gain of electrons by another substance. These processes cannot occur independently; they always happen together in equivalent amounts. This means that for every atom or molecule that is oxidised, there must be another atom or molecule that is reduced.)

Oxidation and reduction can be defined in three different ways:

1) Oxidation is the addition of oxygen to a substance during a chemical reaction. In contrast, reduction is the removal of oxygen from a substance in a chemical reaction. Reduction is the opposite of oxidation.

$$PbO_{(s)} + C_{(s)} \longrightarrow Pb_{(s)} + CO_{(g)}$$

In this reaction, lead oxide is reduced to lead by losing oxygen, whilst carbon is oxidised to carbon monoxide by gaining oxygen.

2) Reduction is the addition of hydrogen to a substance during a chemical reaction, while oxidation is the removal of hydrogen from a substance during a chemical reaction.

$$H_2S_{(g)} + Cl_{2(g)} \longrightarrow S_{(l)} + 2HCl_{(g)}$$

In this example, hydrogen sulphide gas is oxidised to sulphur by losing hydrogen, while chlorine is reduced to HCl by gaining hydrogen.

3) Oxidation is the loss of one or more electrons by a substance, while reduction is the gain of one or more electrons from a substance.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

In this reaction, sodium is oxidised to a sodium ion by losing electrons, while chlorine is reduced to a chloride ion by gaining electrons. The element that loses electrons during the reaction is said to be oxidised, and its oxidation number increases. Conversely, the element that gains electrons during the reaction is said to be reduced, causing its oxidation number to decrease.

Keep in Mind:

Keep OILRIG in mind!

Oxidation Is Loss of Electrons, whereas Reduction Is Gain of Electrons. It is simple and memorable.

6.2 Indication of Oxidation States by Roman Numerals

The oxidation state, or oxidation number, of an atom represents the hypothetical charge it would have in a compound. This charge is positive when electrons are lost (oxidation) and negative when electrons are gained (reduction). It helps us understand how atoms are involved in chemical reactions, particularly in redox (reduction-oxidation) processes. Although it does not always indicate the actual charge on the atom, it is a useful method for tracking electron movement during chemical reactions. For example, in H_2O (water), hydrogen has an oxidation state of +1, while oxygen has an oxidation state of −2.

Roman numerals indicate the oxidation state of an element in a compound when it can exist in more than one oxidation state. For example, in the compound $FeCl_2$, iron (Fe) has an oxidation state of +2, denoted by the Roman numeral II as Fe(II). Similarly, in the compound $FeCl_3$, iron (Fe) has an oxidation state of +3, represented by the Roman numeral III as Fe(III).

Common oxidation numbers are expressed in Roman numerals:

I denotes an oxidation number of +1, II denotes an oxidation number of +2,

III denotes an oxidation number of +3, IV denotes an oxidation number of +4,

V denotes an oxidation number of +5, VI denotes an oxidation number of +6,

VII denotes an oxidation number of +7, and VIII denotes an oxidation number of +8.

Common examples of elements with variable oxidation states include:

1) Lead can exist in two common oxidation states: +1 and +2.

(i) Lead(II) ion: Pb^{2+} (ii) Lead(IV) ion: Pb^{4+}

2) Manganese can exist in four common oxidation states: +2, +4, +6, and +7.

(i) Manganese(II) ion: Mn^{2+} (ii) Manganese(IV) ion: Mn^{4+}

(iii) Manganese(VI) ion: Mn^{6+} (iv) Manganese(VII) ion: Mn^{7+}

Chemistry in Your Body

Cellular respiration is a redox process: glucose is oxidised to release energy, while oxygen is reduced to form water.

6.3 Redox Reactions and the Roles of Reducing and Oxidising Agents

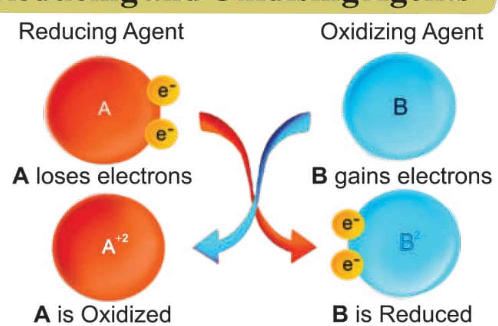


Fig. 6.1: Oxidising and reducing agents

A substance that loses electrons during a chemical reaction is called a reducing agent or reductant, whereas a substance that gains electrons during a chemical reaction is called an oxidising agent or oxidant. The oxidising agent oxidises other substances, increasing their oxidation state, and is itself reduced. Conversely, the reducing agent reduces other substances, decreasing their oxidation state, and is itself oxidised. The total number of electrons gained by the oxidising agent is always equal to the total number of electrons lost by the reducing agent.

Some reactions of oxidising and reducing agents include:

1) Sulphurous acid (H_2SO_3) reacts with nitric oxide (NO) to form sulphuric acid and nitrous oxide (N_2O).

$$H_2SO_3 + 2NO \longrightarrow H_2SO_4 + N_2O$$

In this reaction, sulphurous acid acts as the reducing agent because it (sulphur) loses electrons, while nitric oxide is the oxidising agent as it (nitrogen) gains electrons. The electrons are transferred from sulphur to nitrogen.

2) Magnesium reacts with cold dilute nitric acid to form hydrogen gas.

$$2HNO_3 + Mg \longrightarrow Mg(NO_3)_2 + H_2$$

In this reaction, nitric acid acts as an oxidising agent, while the magnesium atom acts as a reducing agent.

3) Hydrogen sulphide reacts with chlorine to produce hydrogen chloride and sulphur.

$$H_2S + CI_2 \longrightarrow 2\ HC\ell + S$$

In this reaction, hydrogen sulphide acts as reducing agent and chlorine as an oxidising agent.

Keep in Mind

Among the elements, metals are reducing agents, while non-metals are oxidising agents. In the periodic table, alkali metals are strong reducing agents because they readily form positive ions, and halogens are strong oxidising agents because they readily form negative ions.

Among the compounds, potassium dichromate ($K_2Cr_2O_7$) and potassium permanganate ($KMnO_4$) are the most important strong oxidising agents, whereas hydrogen sulphide (H_2S), sulphur dioxide (SO_2), and ferrous sulphate ($FeSO_4$) are some important reducing agents.

6.4 Oxidation Number of Elements in their Free State

When elements are in their pure form or free state, their oxidation number is always zero. This principle applies to all elements, whether they exist as single atoms (such as Na or Mg), diatomic molecules (such as O_2 or N_2), or larger structures. However, it is important to note that the oxidation number of elements can change when they form compounds or undergo chemical reactions, which involve the gain or loss of electrons.

6.5 Derivation of the Formula of Ionic Compounds from the Ionic Charges

The formula of an ionic compound can be derived from the ionic charges of cations and anions, as well as their oxidation numbers. The formula is written so that the charges of the cation and anion sum to zero.

General Steps to Derive the Formula of an Ionic Compound

1) Identify the charges of the cations and anions present in the compound. The charges can frequently be determined from the oxidation states of the elements.
2) Determine the smallest whole-number ratio of cations to anions that produces a neutral compound.
3) Balance the charges of the ions to establish the formula of the compound. The subscripts in the formula are adjusted to ensure that the net charge of the compound remains zero.

Example 1:

In the compound sodium chloride (NaCl), sodium forms the Na^+ ion with a +1 charge, and chlorine forms the Cl^- ion with a −1 charge. The sodium ion (Na^{1+}) has a 1+ charge, so it requires one chloride ion, Cl^-, with a 1 charge to balance the charges. Conversely, the chloride ion (Cl^-) has a −1 charge, meaning it requires one sodium ion (Na^{1+}) with a 1+ charge to achieve charge balance. The charges are already balanced (both 1+ and 1−), so the smallest whole-number ratio between cations and anions is 1:1. Therefore, the formula for sodium chloride can be derived as NaCl.

Test Yourself:

Derive the formula of $Ca(NO_3)$ from the ionic charges. The oxidation state of the calcium ion is 2+, while that of the nitrate ion is 1.

Example 2: Derivation of the formula of Ionic compound (Na_3PO_4)

	Na	PO_4	
Charge of Ion	+1	−3	
No. of Atoms	3	1	Total Charge (on the molecule)
Total Ion Charge	+3	−3	= 0

Table. 6.1: Some common ionic compounds along with cations and anions

Cation	Anion	Formula	Name
Sodium (Na^+)	Chloride (Cl^{1-})	NaCl	Sodium chloride
Magnesium (Mg^{2+})	Oxide (O^{2+})	MgO	Magnesium oxide
Calcium (Ca^{2+})	Chloride (Cl^{1-})	$CaCl_2$	Calcium chloride
Aluminium (Al^{3+})	Oxide (O^{2-})	Al_2O_3	Aluminium oxide
Iron (Fe^{2+})	Oxide (O^{2-})	FeO	Iron(II) oxide
Iron (Fe^{3+})	Oxide (O^{2-})	Fe_2O_3	Iron(III) oxide
Copper (Cu^+)	Oxide (O^{2-})	Cu_2O	Copper(I) oxide
Copper (Cu^{2+})	Oxide (O^{2-})	CuO	Copper(II) oxide
Zinc (Zn^{2+})	Oxide (O^{2-})	ZnO	Zinc oxide
Lead (Pb^{2+})	Oxide (O^{2-})	PbO	Lead(II) oxide
Lead (Pb^{4+})	Oxide (O^{2-})	PbO_2	Lead(IV) oxide

6.6 Principles of Oxidation Numbers in Chemical Species

6.6.1 Oxidation Number of Monatomic Ions

The oxidation state of a monatomic ion is equivalent to its charge. For example, Na^{+} has an oxidation number of 1+, Mg^{2+} has 2+, Al^{3+} has 3+, Cl^{1-} has 1−, O^{2-} has 2−, and N^{3-} has 3−.

6.6.2 Sum of Oxidation Numbers in a Neutral Compound

The sum of the oxidation numbers of all the elements in a neutral compound is equal to zero. For example, in potassium permanganate ($KMnO_4$), the oxidation number of potassium is 1+, manganese is 7+, and that of oxygen is 2−. Hence, the sum of the oxidation numbers is 1(1+) + 1(7+) + 4(2−) = 0.

Self-Assessment

Calculate the oxidation number of sulphur in SO_2, H_2SO_4, and Na_2S. Explain how they differ.

6.6.3 Sum of Oxidation Numbers in an Ion

The sum of the oxidation numbers of all the elements in an ion is equal to the charge on the ion. For example, in the sulphate ion (SO_4^{2-}), the oxidation number of sulphur is 6+, and that of oxygen is 2−. Hence, the sum of oxidation numbers is 1(6+) + 4(2−) = 2−.

6.7 Identification of Redox Reactions

Redox reactions can be identified by observing the changes in colour that occur when acidified aqueous potassium manganate(VII) or aqueous potassium iodide is used. Potassium manganate(VII) is a powerful oxidizing agent and has a purple colour, while potassium iodide is a reducing agent and is colourless in solution. When acidified aqueous potassium manganate(VII) is added to the solution of aqueous potassium iodide, the purple colour of the potassium manganate(VII) disappears. This is because the potassium manganate(VII) is reduced to manganese(II), which is colourless. On the other hand, the colourless potassium iodide solution is oxidised to iodine, which is brown in colour.

Fig. 6.2: Purple colour of potassium manganate(vii) changes to colouless.

$$\underset{\text{Purple}}{2KMnO_4} + 8H_2SO_4 + 10KI \longrightarrow 2MnSO_4 + 6K_2SO_4 + \underset{\text{Brown}}{5I_2} + 8H_2O$$

6.8 Corrosion and Its Prevention

Corrosion is the term applied to metal rust caused by an electrochemical process. Most of you are familiar with the corrosion of iron, which develops as a reddish-brown rust. There are other examples of corrosion as well, such as the black tarnish on silver and the red or green corrosion that may

Fig. 6.3: Rust in iron

develop on copper and brass. Corrosion leads to the deterioration of essential properties in a material. Corrosion is primarily due to redox reactions that occur in the presence of oxygen, moisture, and acidic oxides, such as CO_2. It damages metals and converts them into oxides, which are ultimately converted into hydroxides; for example, iron is converted into Fe_2O_3. nH_2O (rust). Corrosion causes significant damage to metals and other materials worldwide, costing billions of dollars each year.

FUN FACT

One litre of water can produce over 1,200 litres of steam when boiled, illustrating the significant difference in the spacing of gas molecules compar ed to liquids.

6.8.1 Rusting of Iron

The rusting of iron chemically means the electrochemical reaction of iron with oxygen and water to form $Fe(OH)_3$, $Fe(OH)_2$, or even Fe_2O_3. H_2O. Chemically, the corrosion of Iron is written as:

$$4Fe + 3O_2 + H_2O \longrightarrow \underset{\text{(Hydrated Ferric oxide)}}{2Fe_2O_3.nH_2O}$$

MISCONCEPTION Vs REALITY

Myth: Rusting occurs instantly when iron comes into contact with water.
Reality: Rusting is a slow electrochemical process that requires water, oxygen, and often salts to speed up the reaction.

6.8.2 Prevention of corrosion

Several methods have been developed to protect metal from corrosion. The Methods used to prevent corrosion in metal are as follows:

Alloying

The corrosion of metals can be prevented or minimised by reducing their reactivity through alloying. For example, stainless steel, an alloy of iron, chromium, silicon, and nickel, is highly resistant to corrosion. Stainless steel is used for making knives, spoons, forks, utensils, scissors, and surgical instruments.

Fig. 6.4: Pure iron corrodes, whereas stainless steel alloy resists corrosion.

Oil or grease coating:

The corrosion of metal can be prevented or minimised by covering its surface with grease or oil. For example, nuts, bolts, tools, machinery parts, and engine components are coated with grease or oil to protect them from rusting.

Fig. 6.5: Grease or oil coating protects metal from rusting.

Paint coating:

The corrosion of metal can be prevented or minimised by covering its surface with paint. For example, iron bridges, windows, doors, gates, and the bodies of rickshaws, cars, buses, and trucks are coated with paint to protect them from corrosion (rusting).

Fig. 6.6: Paint coating protects metal surfaces from rust and corrosion.

Galvanising

The process by which metal (iron) sheets are coated with a thin layer of zinc to prevent corrosion is called galvanising. It is also known as zinc coating or anode coating. This process can be done by dipping a clean iron sheet into a molten zinc bath. If the zinc coating is damaged by a scratch or dent, the zinc reacts with the air to form a new protective layer. Zinc plating also protects and enhances the appearance of ferrous metals (i.e., iron and steel) as a corrosion barrier, remaining effective even under high humidity and moisture conditions. One of the most common applications of zinc plating is to provide durable and cost-effective rust protection for steel.

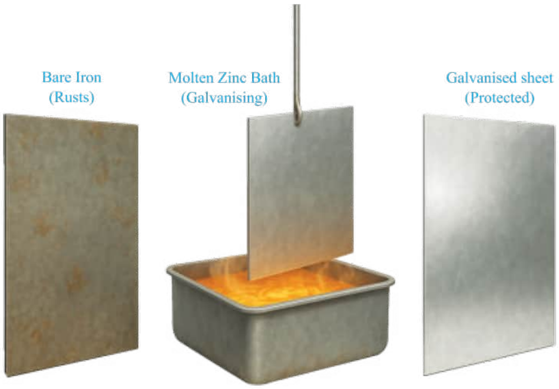


Fig.6.7: Galvanising protects iron by coating it with zinc, preventing rust and corrosion.

Electroplating

A similar protective method is electroplating, in which a thin layer of metal is deposited onto the surface of another metal using an electric current. This process is commonly used to coat metals with nickel, chromium, silver, or gold to enhance corrosion resistance and improve appearance. The object to be coated is placed in an electrolyte solution and acts as the cathode, while the coating metal serves as the anode. When current is applied, metal ions from the solution are deposited onto the object. Electroplating is widely used in industries to protect metals and improve their appearance.

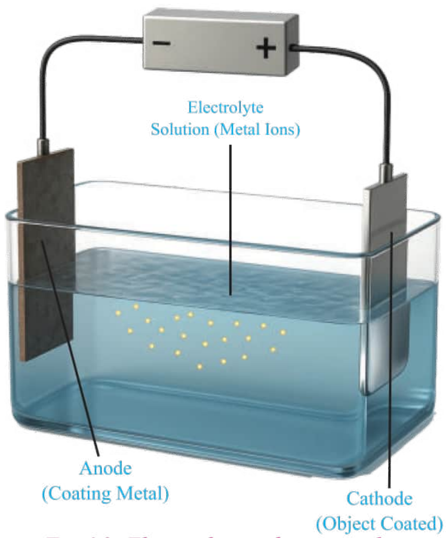


Fig.6.8: Electroplating deposits a thin protective metal coating on another metal through the use of electric current.

CHEMISTRY IN LIFE

Electroplating is used in jewellery making to coat inexpensive metals with gold or silver, giving them a shiny, durable finish.

Sacrificial Coating

In sacrificial protection, a sacrificial anode, made of a more reactive metal, is intentionally attached to the surface of the less reactive metal (cathode) that requires protection. This more reactive metal (sacrificial anode) corrodes preferentially when exposed to oxygen and water, sacrificing itself to prevent corrosion of the less reactive and more valuable metal (cathode). For example, magnesium blocks are used on ships to protect the steel hull from rusting.

Fig.6.9: Magnesium anodes corrode to protect the steel hull from rusting.

AMAZING FACT

If all the rust produced worldwide each year could be collected, it would weigh millions of tonnes, enough to build thousands of ships.

KEY POINTS

1. Electrochemistry is the study of the conversion between chemical energy and electrical energy through redox reactions.
2. Redox reactions involve oxidation and reduction, always occurring together, remembered by OILRIG.
3. Oxidation numbers represent the hypothetical charge on atoms, shown with Roman numerals for variable states (e.g., Fe(II), Fe(III)).
4. Reducing agents lose electrons and become oxidised, while oxidising agents gain electrons and become reduced; metals usually act as reducing agents, whereas non-metals act as oxidising agents.
5. Ionic compound formulae are derived by balancing cation and anion charges so the overall charge is zero, with examples including NaCl, MgO, and PbO_2.
6. Redox reactions can be recognised by colour changes, such as purple acidified $KMnO_4$ becoming colourless when reduced, while colourless KI is oxidised to brown iodine.
7. Corrosion is the electrochemical degradation of metals, such as the rusting of iron into $Fe(OH)_2$, $Fe(OH)_3$, or $Fe_2O_3 \cdot nH_2O$.
8. Corrosion prevention methods include alloying (such as stainless steel), coating with oil or grease, painting, galvanising with zinc, electroplating with Ni, Cr, Ag, or Au, and sacrificial protection.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) In a redox reaction, the process that involves the gain of electrons is:
- (a) oxidation
- (b) reduction
- (c) both oxidation and reduction
- (d) neither

ii) The oxidation state of lead in PbO_2 is:
- (a) 1+
- (b) 2+
- (c) 3+
- (d) 4+

iii) The role of a reducing agent in a chemical reaction is:
- (a) gains electrons
- (b) loses electrons
- (c) oxidises other substances
- (d) reduces other substances

iv) The oxidation number of sodium in its free state is:
- (a) 1
- (b) 0
- (c) 1+
- (d) 2+

v) The sum of oxidation numbers in a neutral compound equals:
- (a) 1
- (b) 0
- (c) 1+
- (d) 2+

vi) The compound that acts as a strong oxidising agent is:
- (a) hydrogen sulphide (H_2S)
- (b) potassium dichromate ($K_2Cr_2O_7$)
- (c) sodium chloride ($NaCl$)
- (d) iron(II) oxide (FeO)

vii) The colour change observed when potassium manganate(VII) reacts in a redox reaction is:
- (a) purple to brown
- (b) brown to purple
- (c) blue to green
- (d) colourless to red

viii) The formula of an ionic compound can be derived by:
- (a) a random combination of ions
- (b) balancing the charges of cations and anions
- (c) adding hydrogen to the compound
- (d) removing oxygen from the compound

ix) The method of corrosion prevention that involves attaching a more reactive metal as an anode to a less reactive metal surface is:
- (a) painting
- (b) galvanising
- (c) sacrificial coating
- (d) oil or grease coating

x) The oxidation state of chlorine in HCl is:

(a) 1+ (b) 1

(c) 0 (d) 2+

B. Restricted Response Questions (RRQs)

i) Define the oxidation number and state the oxidation number of oxygen in carbon dioxide.

ii) Summarise the sum of oxidation numbers in a neutral compound and provide an example to illustrate this concept.

iii) Potassium and chlorine react to form potassium chloride. Is it a redox reaction? If so, why?

iv) Identify the oxidising and reducing agents in the reactions given below by observing the gain and loss of oxygen.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

$$Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(a)} + Cu_{(s)}$$

v) Briefly discuss Roman numerals for oxidation states in compounds. Give examples of elements with variable oxidation states and their Roman numerals.

vi) Describe the rusting process of iron, including the chemical reactions involved.

C. Extended Response Questions (ERQs)

I) Describe the three definitions of oxidation and reduction in electrochemistry. Provide examples for each definition.

ii) Explore methods to prevent corrosion, including alloying, oil or grease coating, paint coating, galvanising, and sacrificial coating.

iii) Explain how the formula of an ionic compound is derived from the charges of its cations and anions. Provide a step-by-step example using the reaction between calcium ions (Ca^{2+}) and phosphate ions (PO_4^{3-}), including the identification of ions and the balancing of charges.

iv) Determine the oxidation state of 'S' in the following species:

(a) H_2S (b) SO_2

(c) SO_3 (d) $Na_2(S_2O_3)_2$

D. Project

Examine corrosion prevention by placing four iron nails in saltwater: one bare, one painted, one coated with oil, and one wrapped with zinc. Observe rusting over a week and compare which method best protects the iron.

UNIT 7

CHEMICAL ENERGETICS

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-62 Explain the idea of a chemical system and its connection with its surroundings, which influences energy transfer during a chemical reaction.

B-63 Differentiate between exothermic and endothermic reactions by giving examples.

B-64 State that thermal energy is called enthalpy change at constant pressure and recognize its sign as negative for exothermic and positive for endothermic reactions.

B-65 Define activation energy as the minimum energy that colliding particles must have for a successful collision.

B-66 Explain that activation energy depends on the reaction pathway, which can be changed using catalysts or enzymes (detailed pathways are not required).

B-67 Draw, label and interpret a reaction pathway diagram for exothermic and endothermic reactions, which includes enthalpy change, activation energy (uncatalyzed and catalyzed), reactants and products.

B-68 Recognize that bond breaking is an endothermic process and bond making is an exothermic process.

B-69 Explain that enthalpy change is the sum of energies absorbed and released in bond breaking and bond forming.

B-70 Calculate the enthalpy change of a reaction given bond energy values.

Introduction

Thermodynamics is the study of energy, work, heat, temperature, and their interrelations. It also addresses the transformation of energy between objects or substances. Chemical energetics applies thermodynamic principles to understand and predict energy changes in chemical reactions. This includes the absorption or release of heat, the formation and breaking of chemical bonds, and the overall energy balance of a reaction. Additionally, chemical energetics considers the energy stored in the chemical bonds of molecules and the mechanisms by which energy is transferred in chemical processes.

7.1 Chemical System and Surroundings

A system refers to anything (substance or mixture) under observation or experimentation in the laboratory. Everything in the universe that is not part of the system is termed the surroundings. In the context of thermodynamics, the combination of the system and its surroundings is collectively known as the universe.

Universe = System + Surrounding

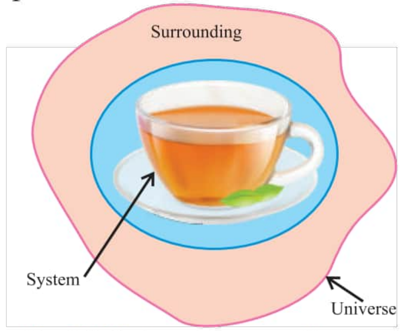


Fig. 7.1: System and surrounding

7.1.1 Types of Chemical Systems

There are three types of chemical systems: open, closed, and isolated.

1) **Open System:** This type of system can exchange energy and matter with its surroundings. Examples include an open reaction flask, a rocket, or an uncovered cup of tea.

2) **Closed System:** A closed system can only exchange energy with its surroundings, not matter (mass). Examples of closed systems include a sealed reaction flask or a gas in a closed container. Chemical energetics commonly studies closed systems, focusing on energy changes.

3) **Isolated System:** An isolated system cannot exchange energy or matter with its surroundings. An example is a thermos flask containing hot green tea, which closely resembles an isolated system.

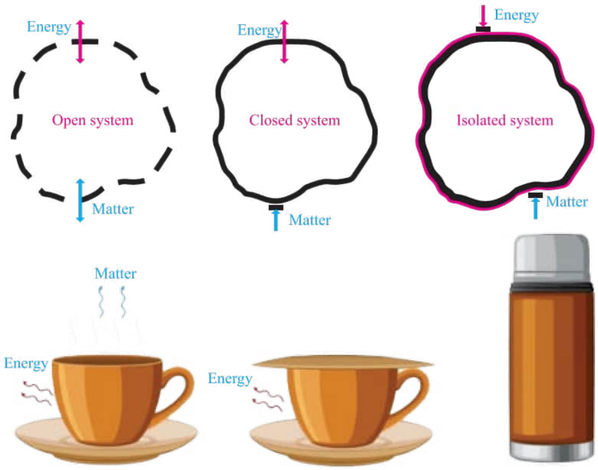


Fig. 7.2: Open, closed and isolated system

7.2 Energy in Chemical Reactions

Chemical reactions can be either exothermic or endothermic.

7.2.1 Exothermic Reactions ($\Delta H < 0$)

Those reactions or processes in which heat is released from the system to the surroundings are known as exothermic reactions or processes (Greek: exō = outside; hence, heat flows out). The combustion of petrol, coal, and wood is an example of an exothermic reaction.

Other examples of exothermic processes include freezing, condensation, and deposition. Some of the exothermic reactions are:

Fig. 7.3: combustion of methane is an exothermic reaction

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow 2H_2O_{(l)} + CO_{2(g)} \qquad \Delta H = -890.3\,kJ/mol$$

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H = -393.51\,kJ/mol$$

The amount of heat released is indicated by ΔH with a negative sign because heat flows from the system to the surroundings. The majority of reactions that occur at room temperature are exothermic in nature.

Do You Know?

The heat produced by burning just 1 gram of fat is almost twice the energy from burning 1 gram of carbohydrate.

7.2.2 Endothermic Reactions (ΔH > 0)

Reactions or processes in which heat is absorbed (from the surroundings to the system) are termed endothermic reactions or processes (Greek: endon = within; thus, heat flows in). The process of photosynthesis in plants is an example of an endothermic reaction. Other examples of endothermic processes include melting, evaporation, and sublimation. Some endothermic reactions are:

Fig. 7.4: Melting of ice is an endothermic process

$$C_{(s)} + H_2O_{(g)} \longrightarrow CO_{(g)} + H_{2(g)} \qquad \Delta H = +131.4\,kJ/mol$$

$$N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)} \qquad \Delta H = +180.51\,kJ/mol$$

The amount of heat absorbed is indicated by ΔH with a positive sign, as heat flows from the surroundings to the system.

Self-Assessment

Is the transformation of a raw egg into a fried egg an endothermic or exothermic process?

CHEMISTRY IN LIFE

Sweating cools the body because the evaporation of water is an endothermic process, which draws heat from the skin and the surroundings.

7.2.3 Enthalpy Change: Thermal Energy in Chemical Reactions

Enthalpy (H) is the total heat content of a system. In thermochemical processes, the change in enthalpy (ΔH) indicates the heat absorbed or released at constant pressure. It is a fundamental concept for understanding endothermic and exothermic reactions, phase changes, and energy transfer in chemistry.

The enthalpy change (ΔH) measures the heat transferred during a chemical reaction and is usually expressed in kilojoules per mole (kJ/mol). It indicates the difference between the energy released when new bonds form in the products and the energy absorbed in breaking the old bonds of the reactants.

FUN FACT

The burning of a candle is both exothermic (releases heat and light) and endothermic (wax melting and vapourising absorbs energy).

7.3 Activation Energy

The minimum amount of energy required to initiate a chemical reaction is known as activation energy, represented by E_a. Its units are joules or kilojoules per mole (kJ/mol).

The reactants are not directly converted into products. They first gain energy to form an activated complex, which then decomposes into products. We can say that activation energy represents an energy barrier (or energy hill) between the reactants and products. The reactant molecules must cross this energy barrier before they can form the products.

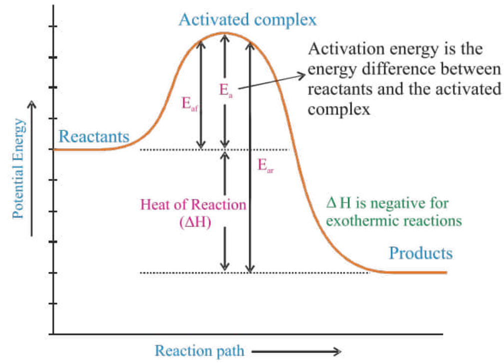


Fig. 7.5: A graph between reaction path and potential energy for exothermic reaction

Crossing this barrier is similar to carrying a ball to the top of a mountain and then rolling it down the other side. However, if the ball fails to reach the peak of the mountain, it will roll back down. Similarly, if activation energy is not provided, the reaction will not start, and the reactants will not be converted into products. This concept can be best understood through the energy diagram.

In exothermic reactions, the potential energy of the reactants is higher than that of the products. This difference is indicated by ΔH.

In endothermic reactions, the potential energy of the reactants is lower than that of the products. This difference is indicated by ΔH. We must provide the activation energy, E_a, to initiate the reaction to form an activated complex.

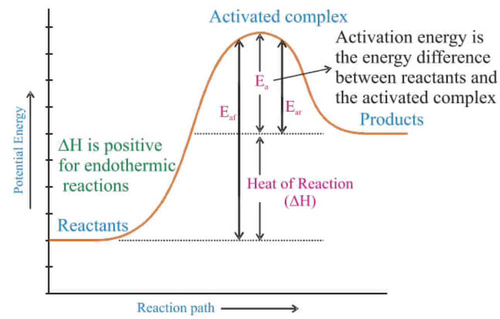


Fig. 7.6: A graph between reaction path and potential energy for endothermic reaction

REAL-WORLD CONNECTIONS

Striking a match requires friction, which provides the activation energy needed to ignite the chemicals on the match head. Without it, the match head will not catch fire.

The activation energy for the forward and reverse reactions varies across different reactions. For exothermic reactions, the activation energy for the forward reaction is lower than that of the reverse reaction because the reactants are at a higher energy level than the products. In endothermic reactions, the activation energy for the forward reaction is higher than that of the reverse reaction because the reactants are at a lower energy level than the products.

Keep in Mind

1. Reactions that are exothermic in the forward direction will be endothermic in the reverse direction.
2. Fast reactions typically have low activation energies, whereas slow reactions have high activation energies.

The Influence of Catalysts and Enzymes on Reaction Pathway

Catalysts (substances that alter the rate of a chemical reaction without being consumed in the process) and enzymes (biological catalysts, typically proteins that accelerate biochemical reactions in living organisms) change the reaction pathway by providing an alternative, more favourable route. This new reaction pathway has a lower activation energy compared to the uncatalysed reaction. Consequently, they facilitate the conversion of reactants into products more efficiently.

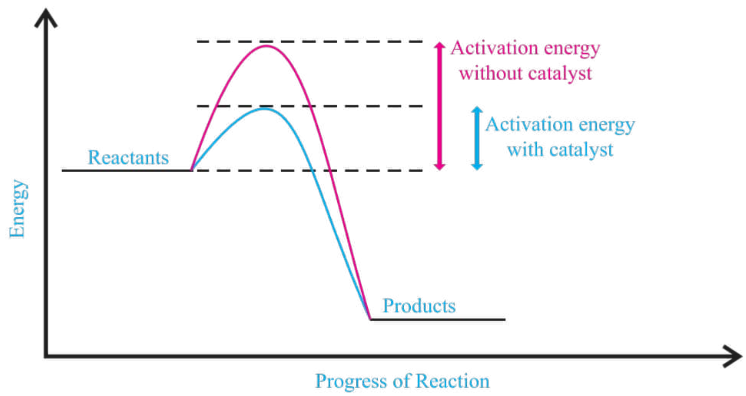


Fig. 7.7: A graph between reaction progress and energy

Keep in Mind:

Catalase is abundant in tissues such as the liver and red blood cells, where it plays a vital role in protecting cells from oxidative damage. The rapid breakdown of hydrogen peroxide by catalase helps maintain cellular health and prevent the accumulation of potentially toxic substances within the body.

$$2H_2O_{2(l)} \longrightarrow 2H_2O_{(l)} + O_{2(g)}$$

7.4 Energy Changes in Chemical Reactions: Bond Breaking and Formation

During chemical reactions, old bonds are broken and new bonds are formed. Bond breakage is an endothermic process, whereas bond formation is an exothermic process.

If the energy needed to break old bonds in the reactants is less than the energy released by new bond formation in the products, then the reaction is exothermic ($\Delta H < 0$). In this case, some of the potential energy is converted into thermal energy. The potential energy of the products is less than that of the reactants; therefore, the products are more stable than the reactants; in other words, the bonds in the products are stronger than those in the reactants.

If the energy required to break old bonds in the reactants is greater than the energy released by new bond formation in the products, then the reaction is endothermic ($\Delta H > 0$). In this case, some of the thermal energy is converted into potential energy. The potential energy of the products is greater than that of the reactants; thus, the products are less stable than the reactants, meaning the bonds in the products are weaker than those in the reactants.

Fig. 7.8: Process of bond breaking and formation

Self-Assessment

Which commonly used unit measures energy changes?

Real-world connections

Exothermic reactions in power plants drive turbines to produce electricity, while endothermic processes, such as photosynthesis, support life on Earth.

7.5 Calculation of Enthalpy change of a reaction

Bond energy is always a positive value because it indicates the amount of energy needed to break one mole of a specific bond in gaseous molecules. Since breaking bonds requires energy, the process is endothermic. Conversely, when new bonds form, energy is released, making the process exothermic. Bond energy is measured in kilojoules per mole (kJ/mol) and helps us estimate the heat of chemical reactions. The enthalpy change of a reaction (ΔH) can be expressed as:

$$\Delta H = \Sigma B.E._{\text{(bonds broken in reactants)}} - \Sigma B.E._{\text{(bonds formed in products)}}$$

Where Σ (sigma) denotes the sum of all bond energies, and B.E. represents bond energy.

The enthalpy change of the reaction, ΔH, is the total energy required to break the existing bonds in the reactants, along with the total energy released during the formation of new bonds in the products. Consider the reaction between hydrogen and chlorine to form HCl. Two moles of HCl gas are produced when one mole of hydrogen gas reacts with one mole of chlorine gas.

$$H_2 + Cl_2 \longrightarrow 2HCl \qquad \Delta H = -432\ kJ/mol$$

This reaction takes place when hydrogen gas reacts with chlorine gas to form hydrogen chloride gas. During this process, one H–H bond and one Cl–Cl bond are broken, while two H–Cl bonds are formed.

Bond energies for H_2, Cl_2, and HCl are given below:

H–H = 436 kJ/mol, Cl–Cl = 242 kJ/mol, and H–Cl = 431 kJ/mol.

Step 1: Energy absorbed to break bonds (Reactants)

Breaking one H–H and one Cl–Cl bond requires: $436 + 242 = 678kJ$

Step 2: Energy released when new bonds form (Products)

Forming two H–Cl bonds releases: $2 \times 431 = 862kJ$

Step 3: Calculate the enthalpy change

$\Delta H = \Sigma$(Bond energies of bonds broken) – Σ(Bond energies of bonds formed)

$\Delta H = 678 - 862 = -184kJ/mol$

Real-Life Application

Energy calculations are essential in designing fuels for rockets, cars, and power stations. Engineers rely on enthalpy data to ensure efficiency and safety.

KEY POINTS

1. Chemical energetics studies the energy changes that occur during chemical reactions.
2. In thermodynamics, a system refers to the substance or mixture under observation, while the surroundings consist of everything else.
3. Chemical systems are classified as open, closed, and isolated.
4. Chemical reactions are exothermic when heat is released ($\Delta H < 0$) and endothermic when heat is absorbed ($\Delta H > 0$).
5. The enthalpy change (ΔH) results from breaking old bonds (which is endothermic) and forming new bonds (which is exothermic).
6. Activation energy (E_a) is the minimum energy required to form an activated complex, acting as a barrier in reactions.
7. Catalysts and enzymes reduce activation energy by altering the reaction pathway, making the conversion of reactants into products more efficient.
8. The enthalpy change (ΔH) of a reaction is calculated by adding the energy needed to break old bonds to the energy released during new bond formation.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The branch of thermodynamics that focuses on energy changes in chemical reactions is:

(a) chemical kinetics (b) chemical energetics
(c) physical chemistry (d) biochemistry

ii) In an endothermic reaction, energy is absorbed:

(a) by the system from the surroundings
(b) by the surroundings from the system
(c) in both directions between the system and the surroundings
(d) no net transfer of energy

iii) The enthalpy change (ΔH) of a reaction is defined as:

(a) heat absorbed or released at constant pressure
(b) variation in temperature of the system
(c) variation in pressure of the system
(d) change in mass of reactants or products

iv) The minimum amount of energy needed to start a chemical reaction is:

(a) enthalpy (b) transition state
(c) energy change (d) activation energy

v) The symbol Δ represents in thermodynamics:

(a) mass of (b) change in
(c) rate of (d) heat stored

vi) In an exothermic reaction, the enthalpy change (ΔH) is always:

(a) zero (b) positive
(c) negative (d) constant

vii) The type of system that exchanges both energy and matter with its surroundings is:

(a) closed system (b) isolated system
(c) open system (d) dynamic system

viii) The role of a catalyst in a chemical reaction is to:

(a) increase activation energy (b) decrease activation energy
(c) produce more reactants (d) inhibit product formation

ix) A negative enthalpy change ($\Delta H < 0$) signifies that the reaction is:

(a) Endothermic (b) Exothermic

(c) Isothermal (d) Non-spontaneous

x) The greater the activation energy:

(a) The lesser the energy needed to start a reaction

(b) The more energy needed to start a reaction

(c) The lesser the energy needed to complete a reaction

(d) The more energy needed to complete a reaction

B. Restricted Response Questions (RRQs)

i) What is the definition of a system in the context of thermodynamics?

ii) Define exothermic reactions and give two examples.

iii) Define activation energy and briefly discuss its role in reaction kinetics.

iv) Briefly explain why breaking bonds is an endothermic process.

v) How can you use the energy from exothermic reactions?

vi) How do bond breaking and bond formation affect energy changes in reactions?

vii) Write a balanced chemical equation for the complete burning of coal.

C. Extended Response Questions (ERQs)

i) Define chemical energetics and explain its importance in understanding chemical reactions, with examples.

ii) Describe open, closed, and isolated systems. How do they differ in matter and energy exchange? Give real-life examples.

iii) How does activation energy differ in exothermic and endothermic reactions? Use examples and energy levels of reactants/products.

iv) How do catalysts and enzymes affect reaction pathways? Explain their effect on activation energy and reaction efficiency.

D. Project

Investigate the enthalpy changes of ammonium chloride and sodium hydroxide. Record the temperature change for each dissolution, classify them as endothermic or exothermic, and present your findings.

UNIT
8

CHEMICAL EQUILIBRIUM

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-72 Recognize that reversible reactions are shown by a symbol ($\rightleftharpoons$) and may not go to completion.

B-73 Describe how changing the physical conditions of a chemical equilibrium system can redirect reversible reactions
(Some examples can include:
a) effect of heat on hydrated compounds
b) addition of water to anhydrous substances, in particular copper (II) sulphate and cobalt (II) chloride.)

B-74 State that reversible reactions can achieve equilibrium in a closed system when the rate of forward and reverse reactions is equal.

Introduction

Chemical equilibrium is the state in a reversible reaction where the rates of the forward and reverse reactions are equal. At this point, the concentrations of reactants and products stay constant over time. This is a dynamic state, meaning that the forward and reverse reactions continue to occur, but at the same rate, so that the overall composition of the reaction remains unchanged. Equilibrium can be affected by changes in temperature, pressure, and concentration.

A good example from everyday life is the equilibrium between carbonated drinks and the carbon dioxide gas above them. In a sealed bottle of soda, carbon dioxide (CO_2) dissolves into the liquid under pressure, while some CO_2 remains in the space above the liquid. An equilibrium is established between the dissolved CO_2 in the drink and the CO_2 gas in the bottle. When the bottle is opened, the pressure drops, disturbing the equilibrium. As a result, CO_2 escapes from the liquid to establish a new equilibrium at the lower pressure, which causes bubbles to form and the drink to lose its fizz gradually.

Fig. 8.1: CO_2 Equilibrium in sealed versus open soda bottles

8.1 Reversible and Irreversible Reactions

8.1.1 Irreversible Reactions

An irreversible reaction proceeds in only one direction, completely converting reactants into products. The reverse reaction may not occur under normal conditions. Such reactions go to completion and are represented by a single-headed arrow (→). For example, the burning of wood, cooking an egg, the digestion of food, the rusting of iron, and the explosion of fireworks.

8.1.2 Reversible Reactions

A reversible reaction is a chemical process that can occur in both the forward and reverse directions. The reactants form products, and the products can also reform the reactants. Such reactions do not go to completion and ultimately attain equilibrium, where the concentrations of reactants and products remain constant. They are represented by double arrows ($\rightleftharpoons$) pointing in opposite directions. Common everyday examples of irreversible reactions include respiration (oxygen and haemoglobin), the formation of carbon dioxide in soda, and the formation of ammonia, etc.

Rate of forward reaction = Rate of backward reaction

8.2 Chemical Equilibrium

The state in which both the forward and reverse reactions occur at the same rate is known as dynamic equilibrium. All chemical equilibria are dynamic. Once the state of chemical equilibrium is established, it will last forever if conditions remain unchanged.

Consider a general reaction where A reacts with B to produce C and D.

$$\underset{\text{(Reactants)}}{A + B} \quad \underset{\text{Reverse reaction}}{\overset{\text{Forward reaction}}{\rightleftharpoons}} \quad \underset{\text{(Products)}}{C + D}$$

At equilibrium, the concentrations of reactants and products remain constant because the forward and reverse reactions occur at equal rates. Although changes continue at the molecular level, the overall amounts of reactants and products do not change.

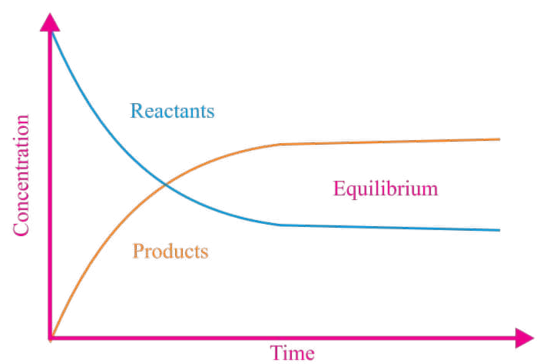


Fig. 8.2: Concentration-time graph

At the start of the reaction, the concentrations of reactants decrease as they are converted into products, while product concentrations increase. As this occurs, the forward reaction slows down and the reverse reaction accelerates. Eventually, both reactions occur at the same rate; this is the equilibrium point.

In a graph depicting concentration versus time, the reactant curves slope downwards while the product curves slope upwards. As equilibrium is achieved, all curves become horizontal, showing that the concentrations remain constant over time.

	Feature of Equilibrium state	Explanation
1	Equilibrium is dynamic.	The reaction has not stopped, but the reverse and forward reactions are still occurring at the same rate.
2	Equilibrium is achieved in a closed system.	A closed system prevents the exchange of matter with its surroundings, achieving equilibrium where both reactants and products can react and recombine.
3	The concentrations of products and reactants are constant at equilibrium.	The rates of product formation and reactant consumption are exactly equal.
4	Equilibrium can be attained from either direction.	The same equilibrium mixture will result under identical conditions, regardless of whether the reaction begins with all reactants, all products, or a combination of both.

8.3 Dynamic Equilibrium in Closed Systems

Reversible reactions can achieve equilibrium (a state of balance) in a closed system when the rates of the forward and backwards reactions are equal. This indicates that the reaction occurs in both directions at the same rate, resulting in no overall change in the concentrations (amounts) of the reactants and products. At equilibrium, the system is balanced, and the concentrations of the reactants and products remain constant over time.

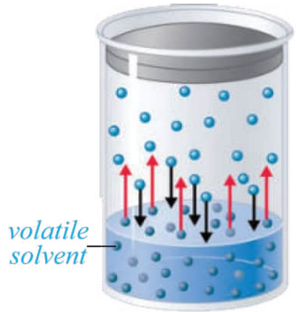


Fig. 8.3: Vapour pressure of a volatile solvent, an example of dynamic equilibrium

For example, consider the vapour pressure of a volatile solvent in a closed container. Initially, the solvent evaporates, increasing the number of vapour molecules above the liquid. As more vapour forms, some molecules start to condense back into the liquid. Eventually, the rate of evaporation equals the rate of condensation, establishing a dynamic equilibrium. At this stage, the amount of liquid and vapour remains constant, even though molecules continue to move between the two phases.

Self-Assessment

Identify at least two essential condition required for equilibrium to be established in a system.

8.4 Effect of Physical Conditions on Equilibrium

8.4.1 Effect of Concentration

Consider a system in equilibrium:

$$\underset{\text{(Reactants)}}{A + B} \rightleftharpoons \underset{\text{(Products)}}{C + D}$$

If the concentration of any reactant (A or B) is increased, the reaction shifts forward to produce more products, while increasing any product (C or D) shifts the reaction backwards to form more reactants. Conversely, decreasing reactants shifts the reaction backwards, and decreasing products shifts it forward to restore equilibrium.

AMAZING FACT

At equilibrium, billions of molecules continue to react every second, but overall concentrations remain constant.

8.4.2 Effect of Pressure

If the pressure of an equilibrium mixture of gases is increased by decreasing the volume at constant temperature, the reaction shifts towards the side with fewer moles of gas molecules to reduce pressure, as in the formation of ammonia, where 4 moles of gas react to form 2 moles.

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

Conversely, if the pressure is reduced by increasing the volume, the reaction shifts towards the side with a greater number of moles of gas, as observed in the dissociation of PCl_5, where 1 mole breaks down into 2 moles.

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

Pressure or volume changes significantly affect equilibrium only when the number of moles of gas differs on each side of the equation.

Chemistry in Your Body

Your lungs maintain equilibrium between oxygen (O_2) and carbon dioxide (CO_2) during breathing, thereby maintaining a stable blood pH.

8.4.3 Effect of Temperature

Changes in concentration, pressure, or volume influence the equilibrium position but do not alter the equilibrium constant, which is only changed by temperature variations. In the formation of nitric oxide, the forward reaction is endothermic, so an increase in temperature favours the formation of products.

$$N_{2(g)} + O_{2(g)} \underset{\text{Exothermic}}{\overset{\text{Endothermic}}{\rightleftharpoons}} 2NO_{(g)} \qquad \Delta H = +180.5\text{kJ/mol}$$

In ammonia formation, the forward reaction is exothermic, so lowering the temperature favours ammonia production.

$$N_{2(g)} + 3H_{2(g)} \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_{3(g)} \qquad \Delta H = -46.2\text{kJ/mol}$$

Increasing the temperature shifts the equilibrium to the left, causing ammonia to decompose.

Interesting Information

An increase in temperature accelerates both exothermic and endothermic reactions as particles gain more kinetic energy and collide more often. However, the rate increases more significantly for endothermic reactions because they absorb heat from the surroundings. This explains why raising the temperature shifts the equilibrium towards the endothermic direction, favouring the reaction that absorbs heat.

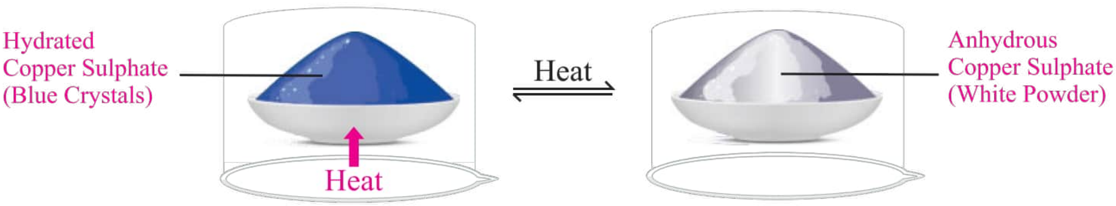
8.4.4 Effect of Heat on Hydrated Compounds

A decrease in temperature favours exothermic reactions, whereas an increase in temperature favours endothermic reactions. For example, when a hydrated compound is heated, the water molecules can be removed, causing the equilibrium to shift towards the formation of the anhydrous form of the compound.

$$\text{Hydrated Compound} + \text{Heat} \rightleftharpoons \text{Anhydrous Compound} + \text{Water}$$

For example, when $CuSO_4.5H_2O$ is heated, the water molecules can be removed, causing the equilibrium to shift towards the formation of $CuSO_4$.

$$CuSO_4.5H_2O_{(s)} + \text{Heat} \xrightarrow{\text{Heat}} CuSO_{4(s)} + 5H_2O_{(g)}$$



This is an endothermic reaction and is favoured by an increase in temperature.

ACITIVITY

Heat the blue $CuSO_4 \cdot 5H_2O$ crystals in a test tube, and observe them change colour to white. Add water again, and the blue colour will return.

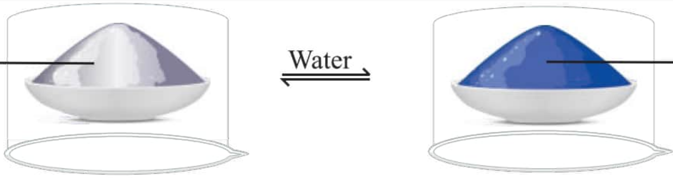
8.4.5 Addition of Water to Anhydrous Substances

Adding water to anhydrous substances can shift the equilibrium towards the formation of the hydrated form, while removing water shifts the equilibrium in the opposite direction. Consider the examples of copper(II) sulfate and cobalt(II) chloride:

Copper(II) Sulphate

Adding water to anhydrous copper(II) sulphate ($CuSO_4$) can shift the equilibrium to the right, favouring the formation of the hydrated form, $CuSO_4.5H_2O$. Conversely, removing water would shift the equilibrium to the left.

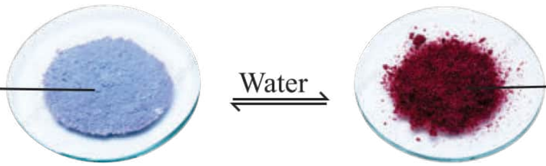
$$CuSO_{4(s)} + 5H_2O_{(l)} \longrightarrow CuSO_4.5H_2O_{(s)} + Heat$$



Cobalt(II) Chloride

Adding water to anhydrous cobalt(II) chloride ($CoCl_2$) can shift the equilibrium to the right, favouring the formation of the hydrated form, $CoCl_2.6H_2O$. Conversely, removing water would shift the equilibrium to the left.

$$CoCl_{2(s)} + 6H_2O_{(l)} \longrightarrow CoCl_2.6H_2O_{(s)} + Heat$$



Safety Note

Copper sulphate and cobalt chloride are hazardous compounds. Always wear safety glasses and gloves when handling these substances. Perform tasks only under the supervision of a teacher. Ensure proper containment and disposal, and follow all safety guidelines.

KEY POINTS

1. In reversible reactions, equilibrium is reached when the rates of the forward and reverse reactions are equal, thereby maintaining a dynamic balance in the concentrations of reactants and products under specific conditions.
2. Reversible reactions are shown by $\rightleftharpoons$, indicating both forward and reverse reactions occur simultaneously, setting them apart from irreversible reactions.
3. The equilibrium point is achieved when the forward reaction rate equals the reverse reaction rate, leading to constant concentrations and parallel lines on a concentration-time graph.
4. All equilibria are dynamic and will continue forever if conditions remain unchanged.
5. Changes in temperature, pressure, and concentration can shift the equilibrium towards the formation of either reactants or products.
6. Exothermic reactions are favoured by a decrease in temperature, while endothermic reactions are favoured by an increase. However, temperature increases the rate of both forward and reverse reactions, because it provides more energy for particles to overcome the activation energy barrier.
7. Heating hydrated salts (such as $CuSO_4 \cdot 5H_2O$) shifts the equilibrium towards the anhydrous form. Adding water reverses this process, reforming the hydrated compound.

8. Adding water promotes hydrated forms, while removal shifts the equilibrium in the opposite direction (e.g., copper (II) sulphate, cobalt (II) chloride).
9. Reversible reactions in closed systems reach dynamic equilibrium when the forward and reverse reaction rates are equal, ensuring that concentrations remain constant over time.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The condition that defines chemical equilibrium is:

(a) forward rate > Reverse rate (b) forward rate < Reverse rate

(c) forward rate = Reverse rate (d) forward rate + Reverse rate = 0

ii) For the reversible reaction

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

Increasing the pressure will shift the equilibrium:

(a) towards reactants (b) towards products

(c) not affected (d) stops the reaction

iii) A reaction is said to be at equilibrium when:

(a) the reaction is complete

(b) product concentration is zero

(c) forward and reverse rates are equal

(d) the forward rate is maximum

iv) The removal of products from a system at equilibrium will:

(a) shift the equilibrium towards the reactants

(b) shift the equilibrium towards the products

(c) keep the equilibrium unchanged

(d) stop the reaction completely

v) A change in physical conditions causes the equilibrium to:

(a) remain unaffected

(b) shift entirely towards reactants

(c) shift entirely towards products

(d) shift towards either reactants or products

vi) The type of equilibrium that exists in chemical reactions at the molecular level is:

(a) static equilibrium (b) temporary equilibrium

(c) dynamic equilibrium (d) mechanical equilibrium

vii) The dissociation of PCl_5 ($PCl_5 \rightleftharpoons PCl_3 + Cl_2$) is favoured by:

(a) low temperature and high pressure
(b) high temperature and low pressure
(c) low temperature and low pressure
(d) high temperature and high pressure

viii) The blue colour of hydrated copper(II) sulphate changes to white on heating because:

(a) water of crystallisation is released
(b) oxygen escapes from the salt
(c) the salt decomposes to copper oxide
(d) sulphur dioxide gas is evolved

ix) When pressure increases in a gaseous system at equilibrium, it:

(a) always shifts right
(b) shifts to the side with more gas molecules
(c) shifts to the side with fewer gas molecules
(d) no effect

x) On hydration, anhydrous cobalt(II) chloride changes colour from:

(a) blue to white
(b) blue to pink
(c) pink to white
(d) pink to blue

B. Restricted Response Questions (RRQs)

i) What is meant by dynamic equilibrium?

ii) Differentiate between static and dynamic equilibrium.

iii) How does increasing the concentration of a reactant affect equilibrium?

iv) What type of reaction is favoured when the temperature is decreased?

v) Explain why the concentration of products remains constant at equilibrium.

C. Extended Response Questions (ERQs)

i) Describe reversible and irreversible reactions, providing examples for each type.

ii) Explain why a closed system is essential for establishing equilibrium, using the example of vapour pressure in a volatile solvent.

iii) Describe the effect of temperature on the equilibrium of hydrated compounds, taking copper(II) sulphate as an example. How does heating affect the position of equilibrium?

UNIT
9

ACIDS AND BASES

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

B-75 Define Bronsted-Lowry acids as proton donors and BronstedLowry bases as proton acceptors.

B-76 Recognizethat aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions.

B-77 Define a strong acid and base as an acid or base that completely dissociates in aqueous solution and weak acid and base that partially dissociates in aqueous solution. (Some examples include: Student writing symbol equations to show these for hydrochloric acid, sulphuric acid, nitric acid, and ethanoic acid.)

B-78 Formulate dissociation equation for an acid or base in aqueous solution.

B-79 Recognize that bases are oxides or hydroxides of metals and that alkalis are water-soluble bases.

B-80 Describe the characteristic properties of acids in terms of their reactions with metals, bases and carbonates.

B-81 Identify the characteristic properties of bases in terms of their reactions with acids and ammonium salts.

Introduction

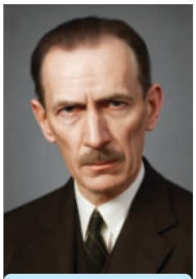
Acids, bases, and salts are essential chemical compounds with numerous applications in our daily lives and various industries. Acids, such as acetic acid in vinegar, can be found in foods and cleaning products. Citric and ascorbic acids are present in citrus fruits, while malic acid is found in apples. Carbonated drinks contain carbonic acid, and sulphuric acid is used in batteries, fertilisers, and explosives.

Conversely, sodium hydroxide, a base, is used in the production of soaps and detergents. Ammonia, another base, is present in household cleaners, while magnesium hydroxide is used as an antacid.

Interesting Information

Citrus fruits such as lemons, limes, and oranges have a sour taste due to the presence of natural acids like citric acid and ascorbic acid (also known as vitamin C). These acids give citrus fruits their characteristic tang, which is balanced with sweetness to create a refreshing and enjoyable flavor.

9.1 Brønsted-Lowry Concept



Bronsted 1879-1747 (Denmark)

The Brønsted-Lowry concept describes the behaviour of acids and bases in terms of proton transfer. It was proposed by Johannes Brønsted and Thomas Martin Lowry in 1923.

According to the Brønsted-Lowry theory, an acid is a substance that donates a proton (H^+). Examples of acids include HCl, HNO_3, H_2SO_4, and CH_3COOH.

$$\underset{\text{Acid}}{HCl_{(g)}} + \underset{\text{Water}}{H_2O_{(l)}} \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+_{(aq)}} + \underset{\text{Chloride ion}}{Cl^-_{(aq)}}$$

Whereas, a base is a substance that can accept a proton (H^+ ion).

Examples include: NH_3, OH^-, and CH_3COO^-.

$$\underset{\text{Ammonia}}{NH_{3(g)}} + \underset{\text{Water}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Ammonium ion}}{NH^+_{4\,(aq)}} + \underset{\text{Hydroxide ion}}{OH^-_{(aq)}}$$

Lowry 1874-1936 (England)

Do You Know?

Lemon juice contains citric acid, which gives it a sour taste, while your stomach produces hydrochloric acid (HCl) to help digest food and kill microbes.

9.2 Aqueous Solutions of Acids

In aqueous solutions, the acids release hydrogen ions (H^+) into the water, making the solution acidic. The release of H^+ ions is a fundamental property of acids and plays an important role in chemical reactions. For example, when sulphuric acid (H_2SO_4) is introduced into water, it dissociates to produce hydrogen ions (H^+) and sulphate ions in the solution.

$$\underset{\text{Sulphuric acid}}{H_2SO_4} + \underset{\text{Water}}{2H_2O} \longrightarrow \underset{\text{Hydronium ion}}{2H_3O^+} + \underset{\text{Sulphate ion}}{SO_4^{2-}}$$

9.2.1 Aqueous Solutions of Alkalis

In aqueous solutions, alkalis (also known as bases) release hydroxide ions (OH^-) into the water, which gives the solution its alkaline nature. This release of OH^- ions is a defining characteristic of alkalis and enables them to neutralise acids. For example, when sodium hydroxide (NaOH), a base, is added to water, it ionizes to produce sodium (Na^+) and hydroxide ions (OH^-) in the solution.

$$\underset{\text{sodium hydroxide}}{NaOH_{(s)}} + \underset{\text{Water}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Sodium ion}}{Na^+_{(aq)}} + \underset{\text{Hydroxide ion}}{OH^-_{(aq)}}$$

9.3 Strong Acids and Strong Bases

The strength of an acid or base is related to the degree of dissociation in water. The extent of dissociation in water distinguishes strong and weak acids/bases. This concept is important to understand the behaviour of substances in aqueous solutions.

9.3.1 Strong and Weak Acids

Strong acids are substances that completely dissociate into ions in aqueous solutions. This results in a high concentration of hydrogen ions in the solution. Examples of strong acids include hydrochloric acid (HCl), sulphuric acid (H_2SO_4), and nitric acid (HNO_3). The dissociation of hydrochloric acid can be represented by the equation:

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

Similarly, the dissociation of sulfuric acid and nitric acid can be represented by:

$$H_2SO_{4(aq)} \longrightarrow 2H^+_{(aq)} + SO^{2-}_{4(aq)}$$

$$HNO_{3(aq)} \longrightarrow H^+_{(aq)} + NO^-_{3(aq)}$$

Weak acids are substances that partially dissociate into ions in aqueous solution. It does not fully dissociate into ions. This results in a low concentration of hydrogen ions in the solution. An example of a weak acid is acetic acid. The dissociation of acetic acid can be represented by the equation:

$$\underset{\text{Acetic acid}}{CH_3COOH} + \underset{\text{Water}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Hydronium ion}}{H_3O^+_{(aq)}} + \underset{\text{Acetate ion}}{CH_3COO^-_{(aq)}}$$

Keep in Mind

The dissociation of a strong acid is represented by a single-headed arrow pointing in one direction, indicating complete dissociation. In contrast, the dissociation of a weak acid is represented by double arrows pointing both ways, indicating an equilibrium between the unionised and ionised forms.

9.3.2 Strong and Weak Bases

Strong bases are substances that completely dissociate into ions in aqueous solution. This results in a high concentration of hydroxide ions in the solution. Examples of strong bases include sodium hydroxide and potassium hydroxide. The dissociation of sodium hydroxide can be represented by the equation:

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

Similarly, the dissociation of potassium hydroxide can be represented by the equation:

$$KOH_{(aq)} + H_2O_{(l)} \longrightarrow K^+_{(aq)} + OH^-_{(aq)}$$

Weak bases are substances that partially dissociate into ions in aqueous solution. This results in a low concentration of hydroxide ions in the solution. An example of a weak base is ammonia. The dissociation of ammonia can be represented by the equation:

$$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

9.4 Differences between Bases and Alkalis

Bases are substances that can accept protons (H^+) and are typically classified as oxides or hydroxides of metals. In short, a base is a proton acceptor. Metal oxides include sodium oxide (Na_2O), calcium oxide (CaO), and copper(II) oxide (CuO). Basic metal oxides react with water to form hydroxides:

$$Na_2O + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)}$$

They neutralise acids to form salt and water.

Real-world connections

Neutralisation is more than just a laboratory experiment; it has practical applications. Farmers add lime $Ca(OH)_2$ to acidic soil to improve crop growth.

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$Ca(OH)_{2(aq)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$$

Bases can be either soluble or insoluble in water. For example, sodium hydroxide dissolves in water, whereas copper(II) oxide does not. Most metal hydroxides are soluble in water and produce hydroxide ions:

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

An alkali is a base that dissolves in water to form hydroxide ions in solution. The solubility of a base determines whether it is classified as an alkali. Sodium hydroxide (NaOH), potassium hydroxide (KOH), and calcium hydroxide ($Ca(OH)_2$) are examples of alkalis:

$$KOH_{(aq)} \longrightarrow K^+_{(aq)} + OH^-_{(aq)}$$

Hydroxide ions (OH^-) are responsible for the basic properties of aqueous solutions. While potassium hydroxide is regarded as an alkali due to its high solubility, calcium hydroxide is only slightly soluble and is generally considered a weak alkali. Conversely, copper(II) hydroxide is insoluble in water and is regarded as a base, not an alkali.

In short, all alkalis are bases, but not all bases are alkalis; only those bases that are soluble in water and produce hydroxide ions in solution are referred to as alkalis.

ACITIVITY

Mix vinegar with baking soda in a beaker. Observe fizzing and the release of carbon dioxide. Hold a lit splint near the bubbles; the flame extinguishes, showing that CO_2 is being produced.

9.5 Characteristic Properties of Acids

9.5.1 Physical Properties

i) They taste sour, similar to lemon juice and vinegar.
ii) They turn blue litmus paper red.
iii) They have a pH value less than 7.
iv) They corrode metals and other materials.
v) They are soluble in water.
vi) They conduct electricity in aqueous solution due to the presence of H^+ ions.

9.5.2 Chemical Properties

i) They react with reactive metals (like sodium, potassium, zinc or magnesium) to produce salt and hydrogen gas. For example:

$$\underset{\text{Potassium}}{2K_{(s)}} + \underset{\text{Hydrochloric acid}}{2HCl_{(aq)}} \longrightarrow \underset{\text{Potassium Chloride}}{2KCl_{(aq)}} + \underset{\text{Hydrogen}}{H_{2(g)}}$$

Keep in Mind:

The reactions of nitric acid with metals tend to differ from those mentioned here, resulting in the formation of other products.

ii) They react with carbonates to produce salt, water and carbon dioxide.

$$\underset{\text{Calcium carbonate}}{CaCO_{3(s)}} + \underset{\text{Nitric acid}}{2HNO_{3(aq)}} \longrightarrow \underset{\text{Calcium nitrate}}{Ca(NO_3)_{2(aq)}} + \underset{\text{Carbon dioxide}}{CO_{2(g)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

iii) They react with bases to produce salt and water. The reaction between an acid and a base is known as a neutralisation reaction.

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$Ca(OH)_{2(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$$

MISCONCEPTION Vs REALITY

Myth: All acids are dangerous and harmful.

Reality: Many weak acids play essential roles in life. Citric acid gives fruits their tangy flavour, ascorbic acid is Vitamin C, crucial for immunity, and acetic acid is commonly consumed in vinegar.

9.5.3 Uses of Acids

Acids are found in cleaning products, food preservation, batteries, metal processing, and various industrial applications. They are commonly used as laboratory chemicals.

Interesting Information

Soft drinks contain carbonic acid (H_2CO_3), which gives them their fizz. While this mild acid enhances flavour, it also gradually erodes tooth enamel if consumed excessively, explaining why dentists recommend moderation in soda intake.

9.6 Characteristic Properties of Bases

9.6.1 Physical Properties

i) They have a bitter taste.

ii) They feel slippery to touch.

iii) Strong bases can be corrosive.

iv) They turn red litmus paper blue.

v) They have a pH value greater than 7.

vi) They conduct electricity in solution due to the presence of OH^- ions.

9.6.2 Chemical Properties

i) They react with metals to produce salt and hydrogen gas.

$$\underset{\text{Zinc}}{Zn_{(s)}} + \underset{\text{Sodium Hydroxide}}{2NaOH_{(aq)}} \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_{2(aq)}} + \underset{\text{Hydrogen gas}}{H_{2(g)}}$$

ii) They react with non-metal oxides to produce salt and water.

$$\underset{\text{Carbon dioxide}}{CO_{2(g)}} + \underset{\text{Calcium Hydroxide}}{Ca(OH)_{2(aq)}} \longrightarrow \underset{\text{Calcium carbonate}}{CaCO_{3(aq)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

iii) They react with acids to produce salt and water.

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

iv) They react with ammonium salts to produce salt, water, and ammonia gas.

$$\underset{\text{Sodiumhy droxide}}{NaOH_{(aq)}} + \underset{\text{Ammonium chloride}}{NH_4Cl_{(aq)}} \longrightarrow \underset{\text{Sodium chloride}}{NaCl_{(aq)}} + \underset{\text{Water}}{H_2O_{(l)}} + \underset{\text{Ammonia}}{NH_{3(g)}}$$

MISCONCEPTION Vs REALITY

Myth: All bases are safe because they neutralise acids.

Reality: Strong bases, such as sodium hydroxide, are highly corrosive and can cause serious burns. Only weak bases, such as baking soda, are safe to use at home.

$$Ca(OH)_{2(s)} + (NH_4)_2SO_{4(aq)} \longrightarrow CaSO_{4(aq)} + 2H_2O_{(l)} + 2NH_{3(g)}$$

Calcium hydroxide	Ammonium sulphate	Calcium sulphate	Water	Ammonia

9.6.3 Uses of Bases

Bases are found in cleaning agents, soap and detergent production, paper manufacturing, water treatment, and various industrial processes. They are commonly used as laboratory chemicals and play an important role in neutralising acids.

CHEMISTRY IN LIFE

Magnesium hydroxide, a weak alkali, is used in "milk of magnesia" to relieve heartburn and indigestion. It neutralises excess stomach acid, providing quick relief.

ACITIVITY

Restoring shine to copper coins with lemon juice.

Rub lemon juice onto an old copper coin and observe it over time.

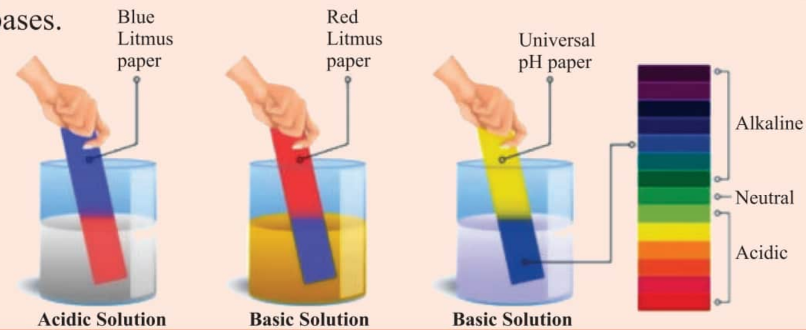
What is the lesson behind the coin's restoration of shine?

ACITIVITY

Objective: The goals are to differentiate between acids and bases using indicators, understand pH, and observe color changes indicating acidity or basicity.

Materials: Droppers or pipettes, small containers or cups, red and blue litmus paper, universal pH indicator paper, and samples of common household substances (lemon juice, vinegar, baking soda solution, soap solution, etc.)

Procedure: Label containers for each substance. Dip red litmus paper into each substance and observe color changes. Repeat the process with blue litmus paper. Place a small amount of each substance on universal pH indicator paper and observe the color change using a dropper or pipette. Record observations and classify substances as acidic, basic, or neutral based on indicators. Discuss results and relate them to the properties of acids and bases.



KEY POINTS

1. The Brønsted–Lowry concept defines acids as proton donors and bases as proton acceptors, a theory proposed by Brønsted and Lowry in 1923.
2. Acids release hydrogen ions (H^+) in aqueous solutions; examples include HCl, HNO_3, and H_2SO_4.
3. Bases release hydroxide ions (OH^-) in aqueous solutions; examples include NH_3, NaOH, and KOH.
4. Strong acids completely dissociate into ions in water, such as Hcl, H_2SO_4, and HNO_3, while weak acids dissociate only partially, as in the case of CH_3COOH.
5. Strong bases like NaOH and KOH also dissociate completely in water, whereas weak bases such as NH_3 dissociate only partially.
6. Alkalis are bases that are soluble in water; while all alkalis are bases, not all bases are alkalis.
7. Acids typically have a sour taste, are corrosive, react with metals, turn blue litmus paper red, and conduct electricity in solution.
8. Bases have a bitter taste, a slippery feel, react with acids, turn red litmus paper blue, and conduct electricity in aqueous solution.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) According to the Brønsted-Lowry theory, an acid is a substance that:
   (a) accepts protons (b) donate s hydroxide ions
   (c) donates protons (d) accepts electrons

ii) The ion responsible for the acidic properties of a solution is:
   (a) OH^- (b) Cl^-
   (c) H^+ (d) Na^+

iii) The best description of an alkali is a:
   (a) base that is insoluble in water
   (b) salt that reacts with water
   (c) base that dissolves in water to produce OH^- ions
   (d) base that dissolves in water to produce H^+ ions

iv) A strong acid among the following is:
   (a) CH_3COOH (b) H_2CO_3
   (c) HCl (d) NH_4OH

v) A neutralisation reaction between an acid and a base typically produces:

(a) acid and gas (b) salt and water

(c) base and oxygen (d) hydrogen gas only

vi) The following is **NOT** a property of acids:

(a) sour taste

(b) react with metals to produce hydrogen gas

(c) turn red litmus paper blue

(d) pH less than 7

vii) When an acid reacts with a carbonate, it produces:

(a) salt and hydrogen gas (b) salt, water, and carbon dioxide

(c) only water (d) salt and oxygen

viii) Sodium hydroxide (NaOH) is:

(a) weak base (b) insoluble base

(c) strong alkali (d) weak acid

ix) The gas released when an acid reacts with a metal is:

(a) oxygen (b) carbon dioxide

(c) hydrogen (d) nitrogen

x) The best description of a neutralisation reaction is:

(a) acid reacts with metal to form hydrogen gas

(b) base reacts with water to produce hydroxide ions

(c) acid reacts with base to form salt and water

(d) acid reacts with carbonate to produce carbon dioxide

B. Restricted Response Questions (RRQs)

i) Which ion do alkalis release in aqueous solutions?

ii) What is the pH value of a neutral solution?

iii) What is formed when an acid reacts with a carbonate?

iv) State one difference between a base and an alkali.

v) Why are strong acids good conductors of electricity in aqueous solutions?

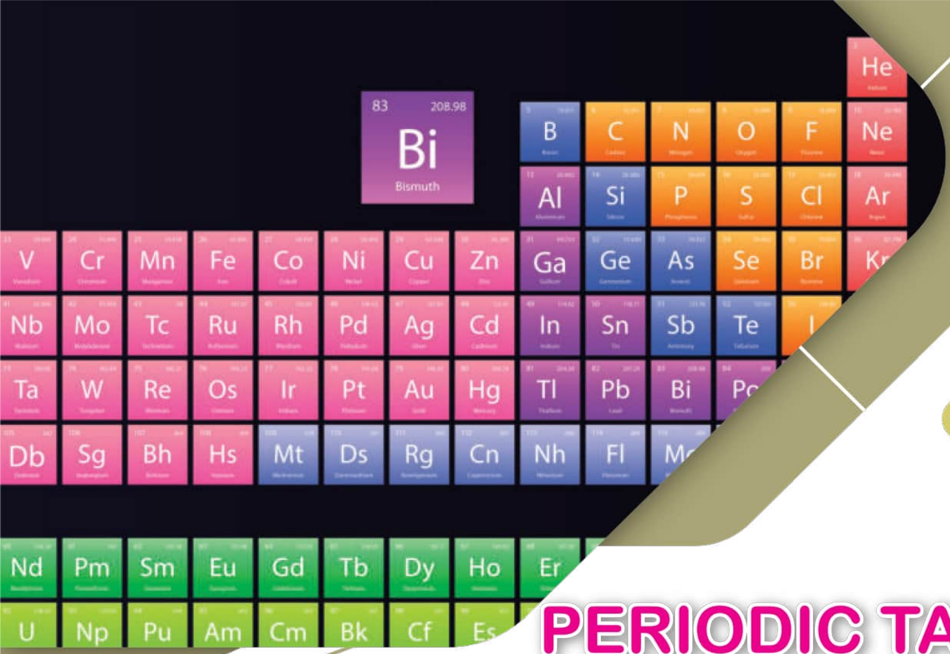
C. Extended Response Questions (ERQs)

i) Explain the Brønsted-Lowry concept of acids and bases. Give one example of each.

ii) Differentiate between strong and weak acids with suitable reactions.

iii) State and explain three physical properties and two chemical properties of acids.

iv) With the help of chemical equations, explain how acids react with metals, carbonates, and alkalis.



UNIT 10

PERIODIC TABLE AND GROUP PROPERTIES

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

C-01 Define the Periodic Table as an arrangement of elements in periods and groups, in order of increasing proton number/atomic number (Note: Use and explain in the Periodic Table group numbers 1-18 and I-VIII).

C-02 Identify the group, period, or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced) .

C-03 Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells .

C-04 Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration .

C-05 Identify trends in group and period, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity and density .

C-06 Use terms alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides and actinides in reference to the Periodic Table.

C-07 Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity .

C-08 Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties.

C-09 Define Group-IA Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.

C-10 Predict properties of other elements in group IA, given information about the elements.

C-11 Predict properties of elements in group-IA in order of reactivity, given relevant information.

C-12 Define group VIIA halogens as diatomic non-metals with general trends limited to increasing density, and decreasing reactivity.

C-13 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.

C-14 Explain the displacement reactions of halogens with other halide ions and also as reducing agents.

C-15 Predict the properties of elements in group-VIIA, given information about the elements.

C-16 Analyse the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths.

C-17 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 examples include catalysts being used in the Haber process, catalytic converters, Contact process and manufacturing of vegetable ghee.)

C-18 Define the Group-VIIIA noble gases as unreactive, monatomic gases.

C-19 Explain noble gases in terms of electronic configuration.

C-20 Compare the general physical properties of metals and nonmetals.
(Specifically in terms of:
a) thermal conductivity
b) electrical conductivity
c) malleability and ductility
d) melting points and boiling points.)

Introduction

The periodic table is a systematic arrangement of elements that group them according to their chemical and physical properties. Elements in the same group exhibit similar characteristics due to their identical valence electron configurations. This organisation not only simplifies the study of elements but also allows us to predict the behaviour of unfamiliar ones based on their position in the periodic table.

10.1 Groups and Periods in the Periodic Table

10.1.1 Groups

The columns of elements in the periodic table are known as groups. There are eighteen groups in the modern periodic table. They contain the same number of electrons in the valence shell (outermost shell). For example, the elements in group 1 (IA) (alkali metals) have one valence electron, whereas those in group 17 (VIIA) have seven valence electrons. Elements within the same group have similar outer electronic configurations and are grouped together. The atomic number of elements increases as you move down the group.

Alkali Metal	Nonmetal
Alkaline Earth	Halogen
Transition Metal	Noble Gas
Basic Metal	Lanthanide
Semimetal	Actinide

6	Atomic Number
C	Symbol
12.011	Atomic Mass
Carbon	Name

1 H 1.008 Hydrogen																	2 He 4.001 Helium
3 Li 6.94 Lithium	4 Be 9.013 Beryllium											5 B 10.81 Boron	6 12.011 Carbon	7 N 14.007 Nitrogen	8 O 15.999 Oxygen	9 F 18.998 Fluorine	10 Ne 20.18 Neon
11 Na 22.98 Sodium	12 Mg 24.32 Magnesium											13 Al 26.98 Aluminium	14 Si 28.085 Silicon	15 P 30.97 Phosphorus	16 S 32.06 Sulfur	17 Cl 35.45 Chlorine	18 Ar 39.95 Argon
19 K 39.10 Potassium	20 Ca 40.08 Calcium	21 Sc 44.96 Scandium	22 Ti 47.87 Titanium	23 V 50.94 Vanadium	24 Cr 51.996 Chromium	25 Mn 54.94 Manganese	26 Fe 55.84 Iron	27 Co 58.93 Cobalt	28 Ni 58.69 Nickel	29 Cu 63.55 Copper	30 Zn 65.38 Zinc	31 Ga 69.72 Gallium	32 Ge 72.63 Germanium	33 As 74.92 Arsenic	34 Se 74.92 Selenium	35 Br 79.904 Bromine	36 Kr 83.80 Krypton
37 Rb 85.47 Rubidium	38 Sr 87.62 Strontium	39 Y 88.90 Yttrium	40 Zr 91.22 Zirconium	41 Nb 92.90 Niobium	42 Mo 95.95 Molybdenum	43 Tc [98] Technetium	44 Ru 101.07 Ruthenium	45 Rh 102.91 Rhodium	46 Pd 106.42 Palladium	47 Ag 107.87 Silver	48 Cd 112.41 Cadmium	49 In 114.82 Indium	50 Sn 118.71 Tin	51 Sb 121.76 Antimony	52 Te 127.60 Tellurium	53 I 126.90 Iodine	54 Xe 131.29 Xenon
55 Cs 132.91 Caesium	56 Ba 137.33 Barium	57 La 138.91 Lanthanum	72 Hf 178.49 Hafnium	73 Ta 180.95 Tantalum	74 W 183.84 Tungsten	75 Re 186.21 Rhenium	76 Os 190.23 Osmium	77 Ir 192.22 Iridium	78 Pt 195.08 Platinum	79 Au 196.97 Gold	80 Hg 200.59 Mercury	81 Tl 204.38 Thallium	82 Pb 207.2 Lead	83 Bi 208.98 Bismuth	84 Po [209] Polonium	85 At [210] Astatine	86 Rn [222] Radon
87 Fr [223] Francium	88 Ra [226] Radium	89 Ac [227] Actinium	104 Rf [267] Rutherfordium	105 Db [268] Dubnium	106 Sg [269] Seaborgium	107 Bh [270] Bohrium	108 Hs [269] Hassium	109 Mt [278] Meitnerium	110 Ds [281] Darmstadtium	111 Rg [282] Roentgenium	112 Cn [285] Copernicium	113 Uut [286] Ununtrium	114 Fl [289] Flerovium	115 Mc [289] moscovium	116 Lv [293] Livermorium	117 Ts [294] Tennessine	118 Og [294] Oganesson

Lanthanide Series	58 Ce 140.12 Cerium	59 Pr 140.91 Praseodymium	60 Nd 144.24 Neodymium	61 Pm [145] Promethium	62 Sm 150.36 Samarium	63 Eu 151.96 Europium	64 Gd 157.25 Gadolinium	65 Tb 158.93 Terbium	66 Dy 162.50 Dysprosium	67 Ho 164.93 Holmium	68 Er 167.26 Erbium	69 Tm 168.93 Thulium	70 Yb 173.05 Ytterbium	71 Lu 174.97 Lutetium
Actinide Series	90 Th 232.04 Thorium	91 Pa 231.04 Protactinium	92 U 238.03 Uranium	93 Np [237] Neptunium	94 Pu [244] Plutonium	95 Am [243] Americium	96 Cm [247] Curium	97 Bk [247] Berkelium	98 Cf [251] Californium	99 Es [252] Einsteinium	100 Fm [257] Fermium	101 Md [258] Mendelevium	102 No [259] Nobelium	103 Lr [266] Lawrencium

Fig. 10.1: Periodic Table of elements

The group of an element is identified by the number of electrons in the outermost shell, also known as valence electrons, particularly for the representative elements (s- and p-block elements). **For s-block elements**, the group number corresponds directly to the number of electrons in the outermost s subshell. For example, an element with an $\mathbf{ns^1}$ configuration is in Group 1, while one with an $\mathbf{ns^2}$ configuration is in Group 2. **For p-block elements**, the group number is determined by adding the number of electrons in the outermost **s and p** subshells and subsequently adding 10 to this sum to align with the standard 18-group periodic table numbering. For example, an element with an outer configuration of $\mathbf{ns^2np^5}$ has 2+5=7 valence electrons, and adding 10 results in Group 17. **In the case of d-block elements**, the group number is the sum of electrons in the outermost s subshell and the **(n−1)d** subshell. For example, an element with $\mathbf{(n-1)d^5ns^2}$ configuration belongs to Group VIIB.

10.1.2 Periods

The rows of elements in the periodic table are called periods. These are the seven periods in the periodic table. As you move from left to right across a period, the atomic number increases, and the elements generally exhibit a gradual change in their properties.

Table 10.1: Periods of elements

Period	Explanation
1	**Shortest period**, contains only two elements: hydrogen and helium.
2	**Short period**, contains eight elements: lithium to neon.
3	**Short period**, contains eight elements: sodium to argon.
4	**Long period**, contains eighteen elements: potassium to krypton.
5	**Long period**, contains eighteen elements: rubidium to xenon.
6	**Longest period**, contains thirty-two elements: caesium to barium, followed by the lanthanides, hafnium to tantalum, and tungsten to radon.
7	**Longest period**, contains thirty-two elements: francium to radium followed by the actinides, rutherfordium to oganesson.

The period of an element is determined by the highest principal quantum number (or energy level) that appears in its electronic configuration. For example, if the outermost electrons of an element occupy the third shell (n = 3), then the element belongs to Period 3. For instance, the electronic configuration of sodium is $1s^2\ 2s^2\ 2p^6\ 3s^1$. Here, the highest energy level that contains electrons is 3, so sodium is in Period 3.

10.2 Blocks in the Periodic Table

All the elements in the periodic table are divided into four blocks. This classification is based on the filled valence electrons in the subshells (s-, p-, d-, and f-).

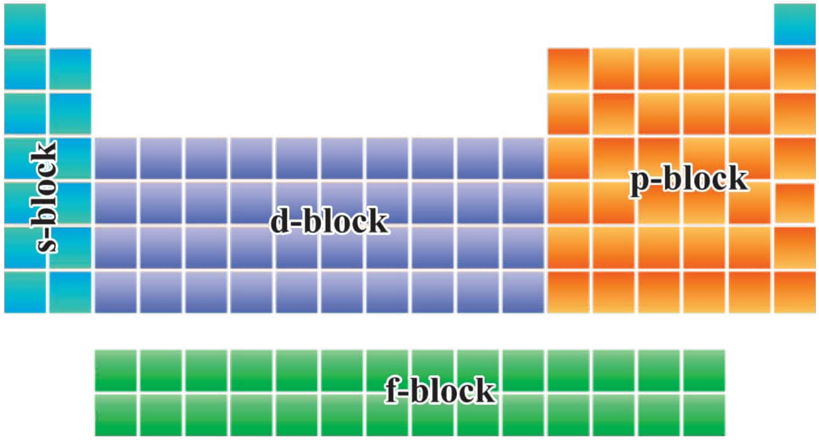


Fig. 10.2: Blocks in the Periodic Table

The block of an element is determined by the type of subshell (s, p, d, or f) into which the last electron enters. If the last electron occupies an s subshell, the element belongs to the s-block. If it enters a p subshell, the element is part of the p-block. If it enters a d subshell, the element is classified as a d-block element, and if it enters an f subshell, the element is categorized as an f-block element. For example, in the case of chlorine, the electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. Since the last electron occupies the 3p subshell, chlorine is a p-block element. Similarly, iron has the configuration $[Ar]\ 4s^2\ 3d^6$; the last electrons fill the 3d subshell, so iron is identified as a d-block element.

10.3 Relationship between Group Number and Charge of Ions in Elements

The relationship between the group number and the charge of ions formed by elements in a group is closely linked to the number of valence electrons. Elements within the same group have the same number of valence electrons. When elements in a group form ions, they either gain or lose electrons to achieve a stable electron configuration similar to that of noble gases.

Elements in Groups 1, 2, and 13 on the left side of the periodic table typically lose electrons to form cations (positive ions) with +1, +2, and +3 charges, respectively. Group 14 elements tend to show +4 or 4 charges. On the right side of the periodic table, elements in Groups 15, 16, and 17 commonly gain electrons to form anions (negative ions) with 3, 2, and 1 charges, respectively. Noble gases in Group 18 possess a complete outer shell and generally do not form ions.

Transition metals in Groups 3 to 12 can show variable oxidation states due to their complex electron configurations. Although lanthanides and actinides are formally part of Group 3, they commonly form M^{3+} ions.

Table 10.2: Typical charges of ions formed by elements in different groups of the periodic table

Group	Typical Charge	Group	Typical Charge
Group 1	+1	Group 16	–2
Group 2	+2	Group 17	–1
Group 13	+3	Group 18	none
Group 14	+4/–4	Transition Metals	variable
Group 15	–3	Lanthanides/Actinides	+3

ACICTIVITY

Exploring Ion Formation and Periodic Trends: Grouping Elements

Objectives: The objectives include introducing the periodic table and valency, understanding the correlation between group number and valency, exploring chemical property similarities within groups, and predicting element positions based on outermost shell electrons.

Materials: Writing board, instruction sheet, periodic table, and cards with symbols of elements and their ions (at least one from each group).

Procedure: The teacher displays the periodic table, highlighting element arrangement and grouping. Students receive cards with element symbols. Eight columns representing groups are drawn on the board. Students discuss ion formation based on element positions. They write ions on their cards, depicting electron movement visually. Cards are displayed on the board, and grouped accordingly. Students explain ion signs, linking to electron gain/loss. The activity ends with a summary of learned concepts.

Results: Students effectively grouped elements on the board based on their ion formation, showcasing an understanding of periodic trends and valency.

Conclusion: The activity facilitated a hands-on exploration of ion formation, fostering an understanding of the periodic table's organization and chemical properties within groups.

10.4 Periodic Trends in Properties of Elements

The physical and chemical characteristics of elements are determined by their arrangement of electrons (electronic configuration), which changes gradually as one moves across the periodic table. Elements within the same group show similar chemical properties because they share the same outer-shell electron arrangement. This indicates that the number of electrons increases as one moves down a group, while the outer shell configuration remains unchanged. Across a period, elements show a gradual change in chemical properties due to progressive changes in their valence shell electron configuration. This recurring pattern of chemical and physical properties at regular intervals, resulting from the periodic recurrence of similar electronic configurations, is known as the periodicity of properties.

ACITIVITY

Online periodic table challenge: Use the internet to find a cross puzzle in the periodic table. Complete with classmates to solve the puzzle in the shortest time possible.

Link: https://www.creative-chemistry.org.uk/funstuff/xword/acids

10.4.1 Atomic Radius

The atomic radius is the average distance from the centre of the nucleus to the outermost electron shell, or half the distance between two identical bonded atoms. When the atomic radius is large, the nucleus is farther from the last electron. Therefore, the attraction between the nucleus and the last electron becomes weaker. This facilitates the atom's ability to give up the electron, thereby increasing its reactivity. Atomic radius is typically expressed in nanometres, angstroms, or picometres.

($1\ nm = 10^{-9}\ m$, $1\ \text{Å} = 10^{-10}\ m$, and $1\ pm = 10^{-12}\ m$).

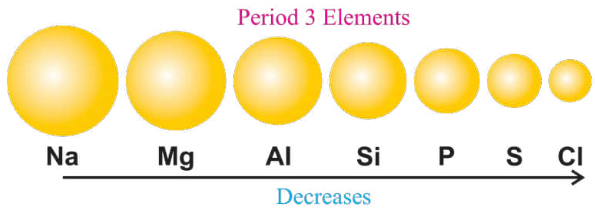
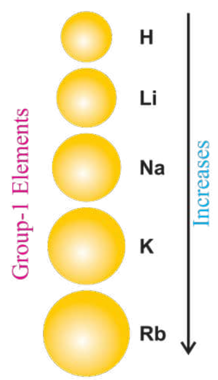


Fig. 10.3: Trends in size of atoms in the periodic table.



The atomic radius increases from the top to the bottom of a group in the periodic table due to an increase in the number of shells. It decreases along the period because of the increase in nuclear charge.

10.4.2 Ionisation Energy

The amount of energy required to remove valence electron from a gaseous atom is called ionisation energy.

$$M_{(g)} \longrightarrow M^{+}_{(g)} + e^{-}$$

For Example:

$$Na_{(g)} \longrightarrow Na^{+}_{(g)} + e^{-} \quad \Delta H = 495.8 kJ/mol$$

$$Mg_{(g)} \longrightarrow Mg^{+}_{(g)} + e^{-} \quad \Delta H = 737.7 kJ/mol$$

The ionisation energy of an element is measured in either kilojoules per mole (kJ/mol) or kilocalories per mole (kcal/mol).

Ionisation energy decreases down the group in the periodic table due to an increase in atomic radius. It increases across a period from left to right because of an increase in nuclear charge.

Keep in Mind:

Electronegativity, electron affinity, and ionisation energy show similar trends in the Periodic Table.

10.4.3 Electron Affinity

Electron affinity is the amount of energy released (or absorbed) when one mole of electrons is added to one mole of neutral gaseous atoms to form negative ions.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

For Example:

$$H_{(g)} + e^- \longrightarrow H^-_{(g)} \qquad \Delta H = -73\ kJ/mol$$

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \qquad \Delta H = -349\ kJ/mol$$

The electron affinity of an element is measured in either kilojoules per mole (kJ/mol) or kilocalories per mole (kcal/mol).

It decreases from top to bottom in a group of the periodic table. This is due to an increase in the number of shells. It increases from left to right in a period of the periodic table. This is due to an increase in nuclear charge.

10.4.4 Electronegativity

Electronegativity is the ability of an atom to attract a shared pair (bond pair) of electrons towards itself within a molecule. Elements with high electronegativity values, like fluorine, attract bonding electrons more strongly than elements like sodium that have low electronegativity values.

In general, electronegativity decreases from top to bottom in a group of the periodic table due to an increase in atomic size. Electronegativity increases from left to right across a period of the periodic table due to the increase in the number of protons in the nucleus (nuclear charge).

Fluorine is the most electronegative element, while caesium is the least electronegative element among those in the periodic table.

ACITIVITY

Use an interactive online periodic table to observe how atomic radius, ionisation energy, and electronegativity change across a period. Record your observations and discuss them in class.
Link: https://ptable.com/#Properties

10.4.5 Metallic Character

The metallic character of an element describes its tendency to lose electrons and form positive ions, whereas the non-metallic character describes its tendency to gain or share electrons to form negative ions or covalent bonds. Metals have low ionisation energies and readily lose valence electrons, resulting in the formation of positively charged cations.

The metallic character of elements decreases along the period due to a decrease in atomic size. As atomic size decreases, the valence electrons become more tightly bound because of an increase in nuclear charge and experience less attraction to the nucleus.

The metallic character of elements increases down the group due to an increase in atomic size. As atomic size increases, the electrons become less tightly bound because of the increase in the number of shells and experience greater attraction to the nucleus.

10.4.6 Reactivity

Reactivity refers to the ability of an element to combine with another element to form a compound. Elements on the left side of the periodic table are metals, whilst those on the right side are nonmetals. Group 1 (IA) elements on the left are known as alkali metals, renowned for their high reactivity, particularly with water. Group 17 (VIIA) elements on the right are known as halogens, which represent highly reactive nonmetals. They react with metals, especially alkali metals, to form ionic compounds, also known as salts.

Reactivity of nonmetals generally increases across a period and decreases down a group. Conversely, metal reactivity typically decreases from left to right across a period. This occurs because the ionisation energy rises as the nuclear charge increases, making it more difficult for metals to lose electrons. However, metal reactivity increases down a group, as atomic size increases, which facilitates the loss of electrons and the formation of cations.

Self-Assessment

1. Why do elements in the same group show similar chemical properties?
2. Explain why lithium reacts less vigorously with water compared to potassium.

10.4.7 Density

Density refers to mass per unit volume. While it may seem obvious that heavier elements are always denser, this is not necessarily the case. The least dense naturally occurring element is hydrogen, whereas iridium and osmium are the densest. Elements created through artificial synthesis, such as meitnerium, are believed to be even denser, although only limited quantities exist for accurate measurement.

Table 10.3: Densities of various elements

Element	H	Ir	Os	Mt	W	Pt	Au	Pb	U	Li
Atomic Number	1	77	76	109	74	78	79	82	92	3
Density (g/cm^3)	0.0000899	22.56	22.59	37.4	19.25	21.45	19.32	11.34	19.05	0.534

Density typically increases down a group and towards the centre of the periodic table, especially among the transition metals. However, elements at the beginning and end of periods tend to have lower densities.

AMAZING FACT

Osmium is the densest naturally occurring element at standard conditions, with a density of about 22.6g/cm^3. The synthetic element meitnerium is predicted to be even denser at around 37.4g/cm^3, but this value is only an estimate, as meitnerium isotopes exist for mere fractions of a second and no bulk sample has been measured experimentally.

10.4.8 Predicting the Properties of Elements Using Periodic Trends

Since periodic trends are consistent within groups and periods, the properties of an element can be predicted reliably based on its position in the periodic table. Elements lower in a group generally have larger atomic radii, lower ionisation energies, reduced electronegativities, stronger metallic character, and increased metallic reactivity. Non-metals higher in a group tend to be more reactive because of their greater ability to gain electrons. Similarly, oxidation states can be predicted: Group 1 elements form +1 ions, Group 2 form +2 ions, Group 17 form –1 ions, and so on. The chemical and physical behaviour of an element can therefore be predicted by comparing it with other members of its group or period.

10.5 Exploring Elemental Groups on the Periodic Table

The periodic table is divided into various groups or families, each with its own specific characteristics.

Alkali Metals: These are Group 1 (IA) elements, which include lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs), and francium (Fr). They are highly reactive and readily form compounds because of their tendency to lose one electron.

Alkaline Earth Metals: These elements are found in Group 2 (IIA) and include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). They are less reactive than alkali metals but still readily lose electrons, forming compounds with a +2 charge.

Halogens: They are Group 17 (VIIA) elements, including fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). These elements are highly reactive nonmetals that typically gain one electron, thus forming salts with metals.

Noble Gases: They are elements of group 18(VIIIA), including helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). These gases are characterised by their low reactivity a complete duplet or octet, which makes them stable and unreactive.

Transition Metals: These occupy the central block (d-block) of the periodic table, encompassing Groups 3 to 12. Examples include iron (Fe), copper (Cu), zinc (Zn), silver (Ag), gold (Au), and numerous others. They show variable oxidation states, enabling the formation of colourful compounds and are essential in various industrial applications.

Lanthanides: This series of 18 elements, located in the f-block, from cerium (Ce) to lutetium (Lu), is comprised of shiny metals with similar chemical properties. They find applications in technologies such as electronics and catalysis.

Actinides: This series of 18 elements is also located in the f-block, below the lanthanides. It includes elements like uranium (U) and plutonium (Pu). These elements are known for their radioactivity and play crucial roles in nuclear technologies and scientific research.

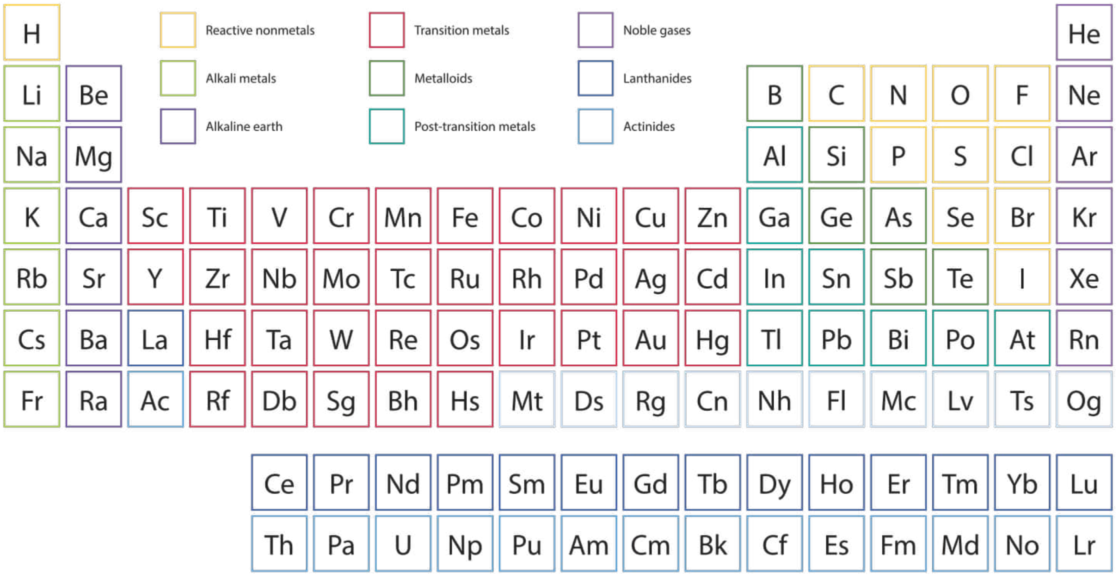


Fig. 10.4: Groups on the periodic table

10.6 Determining the Nature, Position, and Identity of Unknown Elements

When someone provides specific properties (physical state, colour, lustre, density, melting and boiling points, reactivity, specific reactions, valence electrons, ionisation energy, electronegativity, and formation of compounds) of an unknown element, it becomes possible to deduce its nature, possible position in the Periodic Table, and even its identity.

10.6.1 Physical Properties Analysis

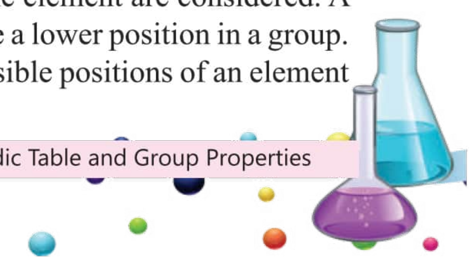
This involves examining whether the element exhibits metallic or non-metallic characteristics based on its physical properties, such as lustre, conductivity, and malleability. If the element has a shiny appearance, good conductivity, and malleability, it suggests a metallic nature. Conversely, if it appears dull, has poor conductivity, and is brittle, it indicates non-metallic characteristics.

10.6.2 Chemical Properties Examination

This step involves investigating the reactivity of the element with water, acids, or other elements, as well as its ionisation energy and electronegativity. A lower ionisation energy suggests a tendency to lose electrons readily, indicating a possible classification as an alkali metal or alkaline earth metal. Higher electronegativity suggests a non-metallic character, possibly placing the element within the halogen or oxygen group.

10.6.3 Position in the Periodic Table Determination

In this step, the atomic radius and group characteristics of the element are considered. A larger atomic radius and lower ionisation energy may indicate a lower position in a group. By examining trends in known group characteristics, the possible positions of an element on the periodic table can be narrowed down.



10.6.4 Comparison with Known Elements

By comparing the observed properties of the unknown element with those of known elements, we can make predictions and narrow down its possible position. Similarities in atomic radius, electronegativity, and chemical reactivity can provide clues about the unknown element's group or period. For example, if the unknown element shares a similar atomic radius, electronegativity, and chemical reactivity with known elements in Group 17 (the halogens), then the possible elements would be fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At).

10.7 Group 1 (IA) Elements

The elements of group 1 (IA) include lithium, sodium, potassium, rubidium, caesium, and francium, and are known as alkali metals. The term alkali is derived from an Arabic word meaning "ashes." This is because many compounds of alkali metals, particularly sodium and potassium, were isolated from the ashes of wood by early chemists. These elements are the most reactive metals in the periodic table. They must be stored in oil (such as kerosene) to prevent reactions with oxygen and water in the air. Their valence shell electronic configuration is "ns^1". They are excellent reducing agents, forming monovalent positive ions (M^+) by losing one valence electron and achieving the stable electronic configuration of noble gases. They show an oxidation state of +1.

Fig. 10.5: Alkali metals

10.7.1 Factors Contributing to the Softness of Alkali Metals in Group IA

The softness of alkali metals (Group 1 elements) in the periodic table can be attributed to their electronic structure and the type of metallic bonding they show. Some important factors that contribute to their characteristic softness include:

Electronic Configuration: Alkali metals have a single electron in their outermost electron shell, which makes them highly reactive. This unpaired electron is relatively far from the nucleus, which weakens the attraction between the outer electron and the positively charged nucleus. This makes it easy for the outer electron to be lost, promoting metallic bonding and resulting in a malleable and soft metal.

Fig. 10.6: Sodium is a soft metal and can be cut by a knife

Metallic Bonding: In alkali metals, the weak attraction between the outer electron and the nucleus permits the easy movement of electrons, thus facilitating metallic bonding. This results in a structure where layers of metal cations can glide past one another with minimal force, making the metal soft and malleable.

Increasing Atomic Size: As you move down Group IA, the alkali metals show an increase in atomic size. The increased atomic size results in a more diffuse outer electron, further weakening the attraction between the outer electrons and the nucleus. This increase in atomic size contributes to the trend of decreasing melting points and increasing softness within the alkali metal group.

Low Melting Points: Alkali metals have lower melting points compared to other metals due to their weak metallic bonding and large atomic size. This means they are soft and can be easily cut with a knife at room temperature.

Increasing Density: As one moves down the group of alkali metals, their density generally increases. This occurs because the increase in atomic size is insufficient to compensate for the increase in mass resulting from the addition of more electrons and protons. The larger atomic volume results in a lower packing density of atoms in the solid state, reducing overall density.

Increasing Reactivity: The reactivity of alkali metals also increases down the group. This is due to their ability to easily lose their outer electron, allowing them to form positive ions. The softness of these metals reflects their highly reactive nature.

10.7.2 Predicted Properties of Other Group IA Elements (Alkali Metals)

Atomic Size: The size of an atom is expected to increase as one moves down a group in the periodic table. This is due to the addition of extra energy levels (shells), which results in larger atomic radii.

Table 10.4: Atomic radii of group IA elements

Element	Li	Na	K	Rb	Cs	Fr
Atomic Radius (pm)	152	186	231	244	262	270

Ionisation Energy: The energy required to remove an electron from an atom is expected to decrease as one moves down a group. This is due to the fact that electrons are located farther from the nucleus, which makes it easier to remove the outermost electron.

Table 10.5: Ionisation energy of group IA elements

Element	Li	Na	K	Rb	Cs
Ionisation Energy (kJ/mol)	520	496	419	403	376

Density: The mass per unit volume of an element is expected to increase as one moves down a group, with the exception of potassium.

Table 10.6: Densities of group IA elements

Element	Li	Na	K	Rb	Cs
Density (g/cm^3)	0.53	0.97	0.86	1.53	1.90

Melting and Boiling Points: The melting and boiling points are expected to decrease as one moves down a group. This is attributed to the weakening forces that hold metal atoms together, which can be linked to the larger atomic size as one proceeds down the group.

Table 10.7: Melting and boiling points of group IA elements

Element	Li	Na	K	Rb	Cs
Melting Point (°C)	180.5	97.8	63.5	39.3	28.4
Boiling Point (°C)	1330	883	760	688	671

Reactivity: The tendency of an element to undergo chemical reactions is expected to increase as one moves down a group. Alkali metals, in particular, are highly reactive; as the outer electrons are further from the nucleus, they become easier to lose, resulting in increased reactivity.

10.7.3 Reactivity Trend in Group 1 (Alkali Metals)

The reactivity of Group 1 elements increases as you move down the group in the periodic table. This occurs because the size of the atoms and the ease with which they can lose their outermost electron both increase as you move down the group.

Lithium (Li) is the least reactive alkali metal because it has the smallest atomic size and the highest ionisation energy among the alkali metals. Its relatively strong hold on its outermost electron makes it less likely to lose that electron during chemical reactions.

Sodium (Na) is more reactive than lithium due to its larger atomic size and lower ionisation energy, which makes it easier for sodium to lose its outermost electron during reactions.

Potassium (K) is even more reactive than sodium because of its larger atomic size and lower ionisation energy, which leads to a more vigorous reaction with water and other substances.

Rubidium (Rb) is more reactive than potassium, showing an increasingly vigorous reaction with water due to its larger atomic size and lower ionisation energy.

Caesium (Cs) is the most reactive alkali metal, reacting explosively with water because of its large atomic size and exceptionally low ionisation energy.

Francium (Fr) is predicted to be the most reactive of the alkali metals due to its position at the bottom of the group; however, its reactivity is difficult to observe because of its scarcity and radioactivity.

The equations for the reactions of alkali metals with water are:

$$2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq)} + H_{2(g)}$$

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

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

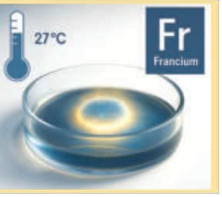
The reactivity of group 1 elements increases as you move down the group. Lithium is the least reactive, while francium is the most reactive.

10.8 Group 17 (VII A) Elements

The group 17 (VIIA) elements, which consist of fluorine, chlorine, bromine, iodine, and astatine, are collectively known as halogens. The word 'halogen' derives from the Greek 'halos' and 'genes', meaning 'salt formers'. This is because all of these elements react directly with metals to produce salts. They have seven electrons in their outermost shells: two electrons in the ns orbital and five electrons in the np orbital. They mostly show a −1 oxidation state.

Interesting Information

Francium is the rarest naturally occurring element, found in only trace amounts on Earth, less than a few grams at any given time. Its extreme rarity and radioactivity make it nearly impossible to study directly.



10.8.1 Physical Properties

i) The halogens exist as diatomic molecules (F_2, Cl_2, Br_2, and I_2), and they are all coloured.

Table 10.8: Physical states and colours of halogens

Element	F_2	Cl_2	Br_2	I_2
Physical States	Gas	Gas	Liquid	Solid
Colours	Pale yellow	Greenish yellow	Reddish brown	Lustrous violet black

ii) They are all poisonous and corrosive.

10.8.2 Trends in Reactivity

Halogens are the most reactive non-metals. Fluorine is the most reactive non-metal known. They can easily accept one electron to complete their octet in chemical reactions. The reactivity of halogens decreases as we move down the group, primarily due to an increase in atomic size and a corresponding decrease in electronegativity. The reactivity of halogens decreases in the following order:

$$F_2 > Cl_2 > Br_2 > I_2$$

Fluorine shows the most excellent reactivity among the halogens due to its small size and high electronegativity.

10.8.3 Trends in density

The density of halogens generally increases from top to bottom within the group. This increase in density is primarily attributed to the rise in atomic mass, rather than a change in atomic size. Although both factors contribute to density, the change in mass has a significantly greater impact.

Table 10.9: Densities of group 17 elements

Halogens	Fluorine	Chlorine	Bromine	Iodine
Density (g/cm^3)	0.0017	0.0032	3.122	4.933

10.8.4 Thermal Stabilities of Hydrogen Halides (HX)

The thermal stability of hydrogen halides primarily depends on the difference in electronegativity between the bonded hydrogen and halogen atoms. Hydrogen fluoride (HF) shows the highest electronegativity difference, resulting in the strongest H–F bond and the greatest thermal stability among all hydrogen halides. Consequently, HF experiences negligible dissociation at high temperatures.

Table 10.10: Thermal stabilities of hydrogen halides

Hydrogen Halides	**Electronegativity Difference**	**Bond Strength (kJ/mol)**	**Thermal Stability**	**Dissociation at High Temperature**
HF	3.98 – 2.20 = 1.78	565	Highest	Negligible
HCl	3.16 – 2.20 = 0.96	432	High	Moderate
HBr	2.96 – 2.20 = 0.76	366	Moderate	Significant
HI	2.66 – 2.20 = 0.46	299	Lowest	Extensive

10.8.5 Halogen Displacement Reactions: Dual Role as Oxidising and Reducing Agents in Redox Reactions

The displacement reactions of halogens with other halide ions involve the transfer of electrons between the halogen atoms. Consider the reaction between chlorine and potassium iodide:

$$Cl_2 + 2KI \longrightarrow 2KCl + I_2$$

In the above reaction, the more reactive halogen (chlorine) displaces the less reactive halogen (iodine) from its salt (KI) solution, leading to the formation of the salt of the more reactive halogen (KCl) and the diatomic molecule of the less reactive halogen (I_2).

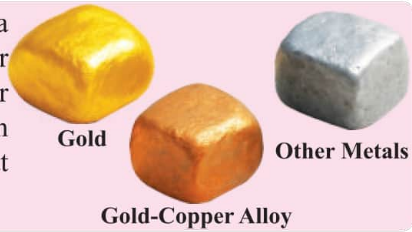
In terms of halogens acting as reducing agents, they undergo redox reactions with metal halides in solution. In this reaction, the more reactive halogen (Cl_2) is reduced by gaining electrons from the less reactive halide ion (I^-). The less reactive halogen is oxidised as it loses electrons. The chlorine atoms are reduced from 0 to −1 oxidation state, while iodine atoms are oxidised from −1 to 0 oxidation state. This dual role of halogens in displacement reactions, acting as both oxidising and reducing agents, is a characteristic feature of redox reactions.

10.9 Transition Elements

The elements in groups 3 to 12 are commonly referred to as transition elements. They are located in the centre of the periodic table, between groups 2 (IIA) and 13 (IIIA). Examples of transition elements include chromium, iron, cobalt, silver, copper, and gold.

AMAZING FACT

Gold is the only metal that naturally shows a yellow hue, whereas most other metals appear silvery or grey. Any yellowish tint in other metals results either from alloying them with gold or from adding specific elements that alter their light-reflecting properties.



10.9.1 General Features of Transition Elements

i) They are all metals, including many common ones such as iron, copper, silver, and gold.

ii) They are good conductors of heat and electricity.

iii) They have high melting and boiling points.

iv) Most transition elements show more than one oxidation state (i.e. variable oxidation state), with a few exceptions. The oxidation states of many transition elements range from +2 to +7; however, their most frequent oxidation states are +2 and +3.


Copper Sulphate


Potassium Dichromate


Cobalt(II) Chloride Hexahydrate


Iron(III)-oxide

Fig. 10.7: Colour of transition metal salts

Table 10.11: The variable oxidation states of transition elements

Element	Atomic Number	Stable Oxidation State	Element	Atomic Number	Stable Oxidation State
Sc	21	+3	Fe	26	+2, +3
Ti	22	+2, +3, +4	Co	27	+2, +3
V	23	+2, +3, +4, +5	Ni	28	+2, +3, +4
Cr	24	+2, +3, +5	Cu	29	+1, +2
Mn	25	+2, +3, +4, +6, +7	Zn	30	+2

v) Most of their compounds are coloured in the solid state or in solution. For example, $CuSO_4$ (blue), $NiSO_4$ (green), $CoCl_3$ (pink).

vi) Most transition elements (Fe, Ni, Pt) and their compounds (V_2O_5, MnO_2, $CuCl_2$) are used as catalysts in many chemical reactions. Some well-known examples of catalysts are listed in Table 10.14:

Table 10.12: Catalysts and their uses

S/No.	Catalyst	Uses
1	Fe	Finely divided iron is used in the Haber - Bosch process for making ammonia.
2	Ni	Raney nickel is used in the hydrogenation of vegetable oil to form ghee.
3	Cu	Copper is used in the oxidation of alcohols to produce aldehydes.
4	Pt	Platinum is formerly used in the Contact process for producing sulphuric acid (H_2SO_4), which involves converting sulphur dioxide (SO_2) to sulphur trioxide (SO_3).
5	V_2O_5	V_2O_5 is used in the oxidation of SO_2 to SO_3 needed for the production of H_2SO_4 in the Contact process.
6	MnO_2	MnO_2 is used for decomposition of $KClO_3$ to produce O_2 gas.

10.10 Group 18 (VIIIA) Noble Gases

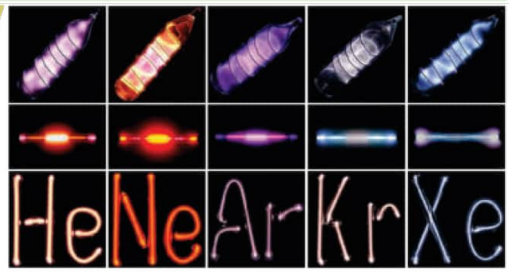


Fig. 10.8: Noble gases

The elements in Group 18 of the periodic table are called noble gases. These include helium, neon, argon, krypton, xenon, and radon. They occur naturally in the atmosphere in very small quantities and are characterised by their extremely low chemical reactivity.

Noble gases are unreactive, monatomic elements. The term monatomic describes elements that exist as single atoms rather than as molecules. Because they do not readily combine with other elements, they are also known as inert gases. Their chemical inactivity is one of their most essential and distinctive characteristics.

10.10.1 Electronic Configuration

The inertness of noble gases is due to their electronic configuration. Helium contains two electrons filling its only shell, while all other noble gases—neon, argon, krypton, xenon, and radon—have eight electrons in their outermost shell. This configuration is referred to as a complete octet.

These elements are chemically stable due to their complete outer shells. Since their atoms already have complete sets of valence electrons, they usually do not gain, lose, or share electrons. As a result, they show minimal chemical reactivity under normal conditions.

Do You Know?

Neon lights glow because of the inert nature of noble gases. When electricity passes through a tube filled with neon, its electrons become excited, producing the bright red-orange glow seen in signboards.

10.10.2 Chemical Behaviour

Noble gases are mostly unreactive because of their stability. Helium, neon, and argon are completely inert and do not form any naturally occurring compounds. However, recent scientific research has revealed that krypton and xenon can form a few compounds with highly reactive elements such as oxygen and fluorine. One example is xenon hexafluoride (XeF_6). These compounds are very rare and can only be produced under specific conditions.

CHEMISTRY IN LIFE

Helium is used to fill balloons and airships because it is non-flammable and less dense than air. Its inert nature makes it a safe alternative to hydrogen for aviation and celebrations.

10.11 Difference in Physical Properties between Metals and Non-metals

The differences in physical properties between metals and non-metals arise from variations in their structures and bonding. Metals have a metallic bond, which involves a "sea" of free electrons shared by the metal atoms. Conversely, non-metals typically have covalent bonds, which involve the sharing of electrons between individual atoms.

10.11.1 Thermal Conductivity

Metals are excellent conductors of heat due to the presence of free electrons, which efficiently transfer thermal energy within the metal structure. Non-metals, in contrast, are poor conductors of heat because they lack the free electrons necessary to transfer thermal energy readily.

10.11.2 Electrical Conductivity

Metals are excellent conductors of electricity because their electrons can move freely, carrying an electric charge. Conversely, non-metals are poor conductors of electricity because they hinder the mobility of charges and are unable to carry an electric charge.

10.11.3 Malleability and Ductility

Metals are malleable (hammered into thin sheets) and ductile (drawn into thin wires) because their atoms are arranged in a regular, repeating structure that enables them to slide past one another without breaking. Non-metals, in contrast, are brittle due to their irregular atomic arrangement and weak forces that result in easy breakage under stress.

10.11.4 Melting Points and Boiling Points

Metals have high melting and boiling points because their strong metallic bonds require a lot of energy to break. Conversely, non-metals have low melting and boiling points because their weaker forces of attraction require less energy to overcome.

MISCONCEPTION Vs REALITY

Myth: Noble gases do not react with any element because they are completely inert. **Reality:** Noble gases are generally very stable due to their complete valence shells, but elements like xenon and krypton can form compounds when subjected to extreme temperature and pressure conditions.

KEY POINTS

1. The periodic table is organised into groups (columns) and periods (rows). Elements within the same group have identical numbers of valence electrons, whereas elements in the same period have the same number of energy levels.
2. The atomic number increases down a group because extra electron shells are added.
3. Elements in groups are assigned family names, including Group IA (alkali metals), Group IIA (alkaline earth metals), Group VIA (chalcogens), Group VIIA (halogens), and Group VIIIA (noble gases).
4. Elements are divided into s, p, d, and f blocks based on the type of valence orbitals.
5. The ions formed are determined by valence electrons: Group IA forms +1 ions, Group IIA forms +2 ions, Group VIIA forms –1 ions, and Group VIIIA remains stable and unreactive.
6. The atomic radius increases down a group and becomes smaller across a period.
7. Ionisation energy and electronegativity decrease down a group but increase across a period.
8. The metallic character increases down a group and decreases from left to right across a period.
9. Metals tend to lose electrons, forming positive ions, whereas non-metals gain or share electrons.
10. Alkali metals are very reactive and are stored in oil to prevent reactions with air or water.
11. Alkali metals each have one outer electron (ns^1), which makes them soft, malleable, and have a low melting point. Atomic size and density increase down the group, while ionisation energy and melting points decrease. Reactivity increases from lithium to caesium.
12. Halogens are reactive diatomic non-metals with coloured states. Their reactivity decreases down the group because atomic size increases and electronegativity decreases. Densities increase, while hydrogen halides become less thermally stable.
13. Transition elements are metallic, excellent conductors, and show variable oxidation states. They form coloured compounds and act as catalysts in numerous chemical reactions.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) In the periodic table, the columns of elements are called:

(a) rows (b) periods
(c) groups (d) families

ii) The elements in Group 17 are known as:

(a) alkali metals (b) halogens
(c) noble gases (d) alkaline earth metals

iii) The number of groups present in the modern periodic table is:

(a) 12 (b) 14
(c) 18 (d) 20

iv) Elements in the same group of the periodic table have:

(a) the same atomic mass (b) similar chemical properties
(c) the exact number of neutrons (d) different electron configurations

v) Atomic radius as you move down a group:
(a) increases (b) decreases
(c) remains the same (d) irregularly

vi) Ionisation energy as you move from left to right across a period:
(a) decreases (b) increases
(c) remains constant (d) varies irregularly

vii) The property that transition metals typically show is:
(a) fixed oxidation states (b) variable oxidation states
(c) high reactivity (d) non-metallic characteristics

viii) The group that typically forms anions with a charge of –1 is:
(a) group IA (b) group IIA
(c) group VIIA (d) group IVA

ix) The alkali metal that reacts most explosively with water is:
(a) lithium (b) sodium
(c) potassium (d) caesium

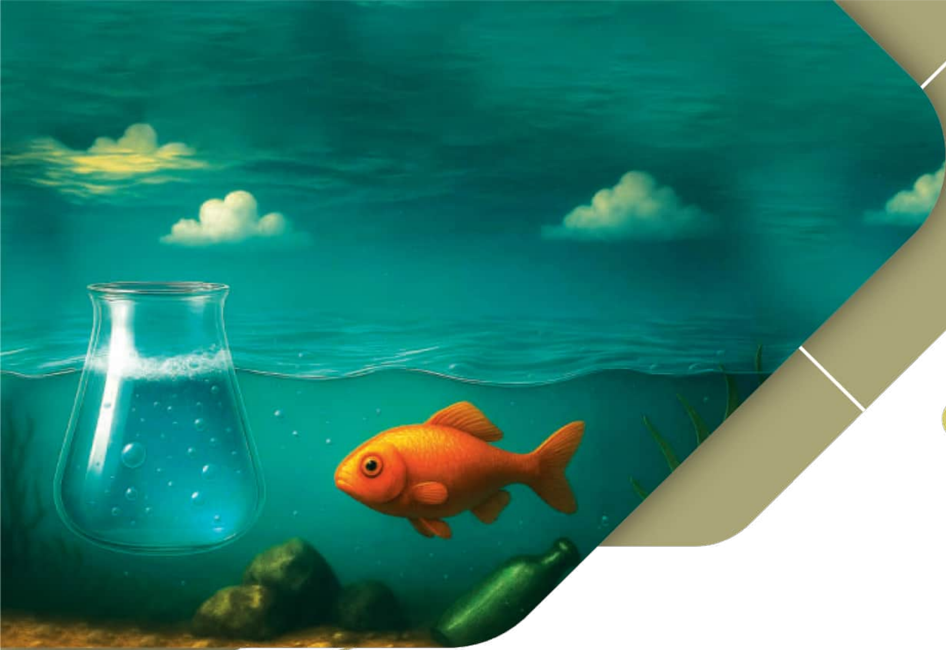
x) Noble gases are unreactive because:
(a) they are non-metals
(b) they have low density
(c) they exist as diatomic molecules
(d) they have complete outer electron shells

B. Restricted Response Questions (RRQs)

i) Why does the atomic radius decrease from left to right across a period?
ii) What are the criteria for classifying an element into the s-, p-, d-, or f-block?
iii) Name three properties that distinguish metals from nonmetals.
iv) Explain why potassium is more reactive than lithium.
v) Describe one chemical property that distinguishes transition metals from alkali metals.
vi) What type of ions do alkali metals form, and why?

C. Extended Response Questions (ERQs)

i) Describe how the electron configuration of an element determines its position in the periodic table.
ii) Define ionisation energy and explain its trend across a period and down a group in the periodic table.
iii) Describe the relationship between group number and the charge of ions formed by elements, including examples from various groups.
iv) Explain the trend in metallic character across a period and down a group in the periodic table.
v) Discuss the electronic configuration of alkali metals and explain how it contributes to their reactivity.
vi) Discuss the general characteristics of transition elements, highlighting their variable oxidation states and catalytic properties.



UNIT 11

ENVIRONMENTAL CHEMISTRY

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

D-01 State that composition of clean, dry air is approximately 78% nitrogen, N_2, 21% oxygen, O_2, and the remainder as a mixture of noble gasses and carbon dioxide, CO_2.

D-02 State the major sources of air pollutants, (Some examples include:
- a) carbon dioxide from the complete combustion of carbon containing fuels
- b) carbon monoxide and particulates from the incomplete combustion of carbon - containing fuels
- c) methane from the decomposition of vegetation and waste gasses from digestion in animals
- d) oxides of nitrogen from car engines
- e) sulphur dioxide from the combustion of fossil fuels which contain sulphur compounds
- f) ground level ozone from reactions of oxides of nitrogen, from car engines, and volatile organic compounds, in presence of light.)

D-03 State the adverse effects of air pollutants , (Some examples include:
- a) carbon dioxide: higher levels of carbon dioxide leading to increased global warming, which leads to climate change
- b) carbon monoxide: toxic gas
- c) particulates: increased risk of respiratory problems and cancer
- d) methane: higher levels of methane leading to increased global warming, which leads to climate change
- e) oxides of nitrogen: acid rain, photochemical smog and respiratory problems
- f) sulphur dioxide: acid rain and haze .)

D-04 Explain how the greenhouse gasses carbon dioxide and methane cause global warming, (Some examples include:

a) the absorption, reflection and emission of thermal energy

b) reducing thermal energy loss to space.)

D-05 Describe the role of sulphur in the formation of acid rain and its impact on the environment.

D-06 Identify the role of ozone in the atmosphere and the harmful effects of ozone depletion.

D-07 Describe the strategies to reduce the effects of major environmental issues. (Some examples include:

a) climate change: planting trees, reduction in emission from livestock farming, decreasing use of fossil fuels, increasing use of hydrogen and renewable energy, e.g. wind, solar

b) acid rain: use of catalytic converters in vehicles, reducing emissions of sulphur dioxide by using low sulphur fuels and flue gas desulphurization with calcium oxide.)

D-08 Describe the role of NO and NO_2 in the formation of acid rain, both directly and through their catalytic role in the oxidation of atmospheric sulphur dioxide.

D-09 Explain how oxides of nitrogen form in car engines and describe their removal by catalytic converters, e.g. $2CO + 2NO \rightarrow 2CO_2 + N_2$

D-10 Define photosynthesis as the reaction between carbon dioxide and water to produce glucose and oxygen in the presence of chlorophyll and using energy from light .

D-12 Identify high risk situations in life including those where long-term exposure to these pollutants can lead to respiratory issues and reduction in quality of life.

D-13 Investigate chemical tests for the presence of water using anhydrous copper (II) sulphate.

D-14 Explain how to test the purity of water using melting point and boiling point.

D-15 Distinguish between distilled water and tap water with their applications in practical chemistry.

D-16 State that water from natural sources may contain useful and harmful substances, (Some examples include:

a) dissolved oxygen b) metal compounds

c) plastics d) sewage

e) harmful microbes f) nitrates from fertilizers

g) phosphates from fertilizers and detergents)

D-17 Recognize that some naturally occurring substances in water are beneficial. (Some examples include:

a) dissolved oxygen for aquatic life

b) some metal compounds provide essential minerals for life.)

D-18 Recognize that some naturally occurring substances in water are potentially harmful. (Some examples include:

a) some metal compounds that are toxic

b) some plastics that harm aquatic life

c) sewage that contains harmful microbes which cause disease

d) nitrates and phosphates that lead to deoxygenation of water and damage to aquatic life; details of the eutrophication process is not required.)

D-19 Explain the treatment of the domestic water supply.
(Some examples of this includes:
a) sedimentation and filtration to remove solids
b) use of carbon to remove tastes and odours
c) chlorination to kill microbes.)

D-20 Describe various water-borne diseases and the steps that can be taken to avoid them.

D-21 Identify the negative effects of water pollutants on life and the ways to avoid them.

D-22 Explain water scarcity as an important issue faced by Pakistan and the ways in which it can be resolved.

D-23 State that urea, ammonium salts and nitrates are used as fertilizers.

Introduction

The atmosphere is the protective layer of gases surrounding the Earth. It sustains life on our planet and shields it from dangerous cosmic rays originating from outer space. It also filters harmful ultraviolet radiation emitted by the sun. Additionally, it plays a vital role in maintaining the Earth's heat balance by absorbing the infrared radiation emitted by the sun and re-emitted by the planet Earth.

However, human activities release pollutants such as carbon monoxide and sulphur dioxide into the atmosphere, threatening air quality and the environment. These pollutants can contribute to air pollution and climate change, and have adverse effects on human health and ecosystems. To preserve the delicate balance vital for life on Earth, it is crucial to engage in environmental conservation and minimise the emission of these harmful substances into the atmosphere.

Water is essential for all living beings. Every plant, animal, and human needs water to survive. It helps us in digestion, keeps us cool, and transports nutrients around the body. Although Earth contains a lot of water, most of it is salty or frozen, leaving only a small amount fresh and drinkable. We must conserve and protect our water to ensure the survival of all living things.

Do You Know?

Without the ozone layer, life on Earth would be impossible. The stratospheric ozone absorbs up to 99% of the harmful ultraviolet radiation from the Sun.

11.1 The Composition of the Atmosphere

The composition of the atmosphere is essential for supporting life on Earth. For example, nitrogen is used by certain bacteria that convert it into useful compounds for plants and by industries to produce ammonia for fertilisers. Oxygen is essential for the survival of living organisms, while carbon dioxide is necessary for plants to carry out photosynthesis. Furthermore, it acts as a carrier of water from the oceans to the land. Its thickness is approximately 500 km above the surface of the Earth.

Table. 11.1: Composition of clean, dry air

Component	Chemical Formula	Percentage by Volume (%)	Significance
Nitrogen	N_2	78.00	A major component of the atmosphere
Oxygen	O_2	21.00	Supports respiration and combustion
Argon	Ar	0.93	Chemically inert; used in lighting and industry
Carbon dioxide	CO_2	0.03	A significant greenhouse gas; essential for photosynthesis
Trace gases	Ne, He, CH_4, Kr, Xe, etc.	~0.01	Present in minute quantities; used in various scientific applications.

11.2 Major Sources of Air Pollutants

Pollutants can be defined as any substance in the air that has a harmful effect on living organisms or their environment. Some major sources of air pollutants are:

11.2.1 Carbon Dioxide (CO_2)

This gas is colourless and typically odourless at low concentrations, but it has a sharp, acidic smell at higher levels. It is soluble in water.

Major Sources

Major sources of CO_2 include the burning fossil fuels such as coal, oil, and natural gas for electricity generation and transportation. These fossil fuels undergo complete combustion with sufficient oxygen, producing CO_2 and H_2O.

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(l)}$$

Fig. 11.1: Methane combustion a source of carbon dioxide emissions

Other sources include industrial processes, deforestation, and the combustion of biomass and waste.

Adverse Effects

Higher levels (more than 65%) contribute to global warming, which in turn causes climate change. Its adverse effects include the melting of polar ice caps and glaciers, leading to rising sea levels and coastal flooding, as well as an increased frequency and intensity of extreme weather events such as heatwaves, droughts, floods, and wildfires.

11.2.2 Carbon Monoxide

Carbon monoxide (CO) is a colourless, odourless, and tasteless gas.

Major Sources

Carbon monoxide is produced from the incomplete combustion of carbon-containing fuels. It is primarily emitted by vehicles, such as cars, trucks, buses, and motorcycles. Industrial operations, including petroleum refining, the paper industry, coal mining, and activities involving electric and blast furnaces, also generate CO. Moreover, forest fires and the incineration of biomass contribute to its production.

Adverse Effects

It is extremely toxic and causes suffocation if inhaled. Exposure of high concentration of CO causes headache, fatigue, unconsciousness and eventually death may occur if such exposure is experienced for a longer time.

CHEMISTRY IN LIFE

Carbon monoxide from car exhaust binds with haemoglobin in your blood, preventing oxygen transport. That is why vehicle pollution can be deadly in congested cities.

11.2.3 Particulates

Particulates (e.g., soot, dust, smoke) are tiny solid or liquid particles suspended in the air.

Major Sources

Particulates are released into the atmosphere from both natural and human-made sources. Natural sources include dust storms, forest fires, volcanic eruptions, and sea spray. Human-made sources primarily involve the burning of fossil fuels in vehicles, power stations, and industrial processes.

Adverse Effects

Particulates irritate the respiratory system and can worsen asthma and bronchitis. Prolonged exposure is associated with lung cancer and cardiovascular diseases. Additionally, particulates diminish visibility by scattering and absorbing light, thereby contributing to haze and compromised air quality.

11.2.4 Methane (CH_4)

Methane is a saturated hydrocarbon. It is the main component of natural gas and is a colourless and odourless gas.

Major Sources

Methane mainly originates from agricultural sources, wetlands, landfills, and flooded rice paddies. Ruminant animals like cattle, sheep, cows, and goats produce methane during digestion. Wetlands, landfills, and rice paddies generate methane through anaerobic decomposition of organic matter (cattle waste) by bacteria.

$$\text{Cattle Waste} \xrightarrow[\text{Organic matter}]{\text{Anaerobic Bacteria}} CO_2 + CH_4$$

Methane can also be emitted by vehicles. Other sources of methane include coal, wood and petroleum.

Adverse Effects

Methane is a powerful greenhouse gas that contributes to global warming and climate change.

11.2.5 Oxides of Nitrogen (NO_x)

The three oxides of nitrogen, N_2O, NO, and NO_2, are important components of the atmosphere. Nitrous oxide (N_2O) is present in relatively high concentrations in the Earth's atmosphere, yet it is not generally regarded as a pollutant. Conversely, nitric oxide (NO) and nitrogen dioxide (NO_2) are more significant pollutants and are often collectively represented as NO_x. NO is a colourless, odourless gas, whereas NO_2 has a reddish-brown colour and a pungent, suffocating odour.

Fig. 11.2: Emission of nitrogen oxide from motor vehicles

Major Sources

The major man-made sources of nitrogen oxides from gasoline and diesel vehicles are a significant contributor to NO_x emissions

$$N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)}$$

$$2NO_{(g)} + O_{2(g)} \longrightarrow 2NO_{2(g)}$$

Adverse Effects

Nitrogen oxides (NO_x) can irritate the airways and lungs, and contribute to the formation of fine particulate matter and ground-level ozone. In the atmosphere, nitrogen oxides combine with rainwater to form acid rain.

$$2NO_{(g)} + O_{2(g)} \longrightarrow 2NO_{2(g)}$$

$$2NO_{2(g)} + H_2O_{(l)} \longrightarrow HNO_{3(aq)} + HNO_{2(aq)}$$

Acid rain can harm aquatic ecosystems, damage forests and crops, and erode buildings and infrastructure.

11.2.6 Oxides of Sulphur (SO_x)

Sulphur dioxide is a colourless, pungent gas, whereas sulphur trioxide is a colourless solid. Both compounds are highly reactive and can irritate or suffocate at higher concentrations.

Major Sources

Sulphur dioxide (SO_2) and sulphur trioxide (SO_3) are formed during the combustion of sulphur-containing fossil fuels in power plants and petroleum industries. The reaction occurs when sulphur compounds in the fuels combine with oxygen, resulting in the formation of SO_2. The sulphur dioxide further reacts with oxygen to produce sulphur trioxide.

Fig. 11.3: Emission of sulphur doxide from power station

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

$$SO_{2(s)} + O_{2(g)} \longrightarrow SO_{3(g)}$$

Adverse Effects

Sulphur dioxide (SO_2) irritates the eyes, nose, throat, and lungs, leading to premature mortality. In the atmosphere, sulphur dioxide and sulphur trioxide mainly react with water vapour to form sulphurous acid and sulphuric acid (H_2SO_4), as major components of acid rain.

$$SO_{2(g)} + H_2O_{(l)} \longrightarrow H_2SO_{3(aq)}$$

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$

AMAZING FACT

One large volcanic eruption can emit more sulphur dioxide in just a few days than all of human industry produces in a year.

Acid rain can also react with fine particulate matter to form haze, which reduces visibility and leads to respiratory issues. Acid rain acidifies soils, rivers, lakes, and streams, harming aquatic life. Additionally, it can damage leaves, plants, and forests, as well as building materials such as steel, paint, cement, limestone, and marble.

11.2.7 Ground Level Ozone (O_3)

Ozone is an allotrope of oxygen. It is a pale-yellow gas, slightly soluble in water, and has a sweetish taste. In the stratosphere, it protects the Earth from harmful UV radiation. Conversely, in the biosphere, ozone acts as a pollutant.

Major Sources

Ground-level ozone is formed from reactions between nitrogen oxides from car engines and volatile organic compounds in the presence of sunlight. These reactions occur naturally but are also influenced by human activities such as the combustion of fossil fuels, which releases NO_x and VOCs into the atmosphere.

Adverse Effects

Ground-level or tropospheric ozone can damage living organisms, including humans, plants, and animals. Exposure to high levels of ground-level ozone can cause eye irritation, throat discomfort, coughing, and asthma, especially with prolonged exposure. Furthermore, it can reduce agricultural productivity. Additionally, it weakens the durability and appearance of paints and fabrics.

Interesting Information

The word "smog" was first introduced in 1905 to describe the combination of fog and smoke that covered London during the Industrial Revolution.

11.3 Role of Ozone in the Atmosphere and Harmful Effects of Ozone Depletion

11.3.1 Role of Ozone in the Atmosphere

Ozone (O_3) occurs in the stratosphere, forming a protective layer about 16–50 km above the Earth's surface.

$$O_2 \xrightarrow{UV} O + O$$

$$O + O_2 \longrightarrow O_3$$

The ozone layer acts as a shield around the Earth, absorbing most of the Sun's harmful rays. This protection prevents genetic mutations, skin cancers, and eye cataracts in humans, while also protecting animals and plants from radiation damage. In this way, ozone plays a vital role in maintaining the balance of life on Earth.

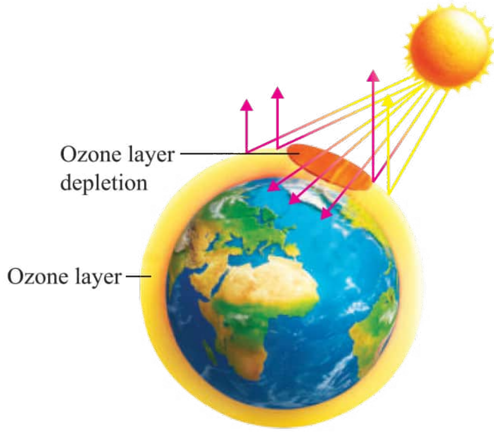


Fig.11.4: Ozone layer depletion

11.3.2 Harmful Effects of Ozone Depletion

A decrease in the amount of ozone in the stratosphere below its normal level is called ozone depletion. As this layer thins, more ultraviolet radiation reaches Earth's surface, causing harmful effects on humans, plants, and the environment.

Effect on Human Health

Depletion of the ozone layer increases the risk of skin cancer, cataracts, and immune system issues, and may lead to respiratory problems due to increased UV exposure.

Effect on Ecosystems

Enhanced UV radiation affects both land and water ecosystems, damaging phytoplankton, disturbing food chains, and altering biochemical cycles.

Effect on Agriculture and Materials

Higher levels of UV radiation can hinder plant growth, reducing agricultural yields. It also damages tyres, rubber, and asphalt, and reduces their durability and strength.

11.4 Greenhouse Effect and Global Warming

11.4.1 Greenhouse Effect

The greenhouse effect is the process by which radiation (thermal infrared radiation) from the Earth is trapped by gases such as carbon dioxide, water vapour, methane, nitrous oxide, and ozone, preventing heat from escaping back into space. Although some of the energy passes back into space, most remains trapped in the atmosphere, causing the Earth to warm. This natural process of maintaining the average temperature of our planet and keeping it warm is known as the greenhouse effect.

11.4.2 Global Warming

Global warming, also referred to as climate change by scientists, is the gradual increase in the average temperature of the Earth's atmosphere resulting from the accumulation of gases that trap heat, preventing it from escaping into space. These gases, known as greenhouse gases, include carbon dioxide, water vapour, methane, nitrous oxide, and ozone. Among them, carbon dioxide and water vapour play a key role in maintaining the Earth's warmth, which is necessary for sustaining life. Excessive amounts of these gases can cause the Earth to become increasingly hotter.

The relative contributions of various greenhouse gases to global warming depend on their concentration in the atmosphere and their ability to absorb and retain heat. Carbon dioxide (CO_2) makes up approximately 76% of the overall greenhouse effect. Methane (CH_4) and nitrous oxide (N_2O) contribute 16% and 6%, respectively, to the total greenhouse effect.

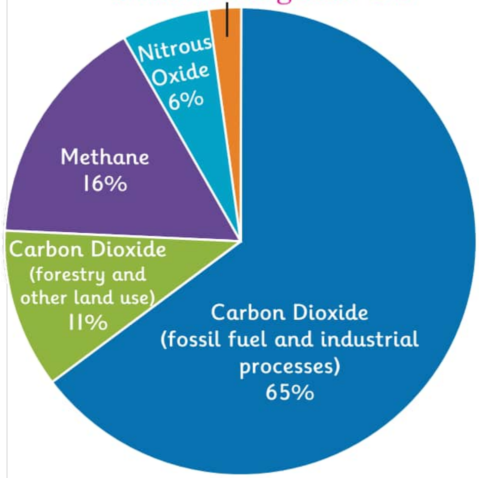


Fig.11.5 : Global greenhouse gas emissions

Real-World Connections

The melting of glaciers in northern Pakistan, especially in the Karakoram and Himalayan ranges, is a clear result of global warming. It leads to Glacial Lake Outburst Floods (GLOFs) that threaten nearby communities, farmland, and infrastructure.

Mechanisms of Global Warming

The mechanism through which these greenhouse gases contribute to global warming is outlined as follows:

A) Absorption, reflection, and emission of thermal energy:

i) Absorption: Greenhouse gases absorb infrared radiation (thermal energy) emitted by the Earth's surface.

ii) Reflection: They then reflect this energy back towards the Earth, preventing it from escaping into space.

iii) Emission: These gases re-emit the absorbed energy in all directions, including back towards the Earth's surface, thereby increasing the overall temperature of the lower atmosphere and the Earth's surface.

B) Reducing Thermal Energy Loss to Space

Greenhouse gases form a barrier that limits the amount of thermal energy escaping into space. By retaining more heat within the Earth's atmosphere, these gases contribute to global warming.

Carbon dioxide is the only greenhouse gas whose contribution is rising rapidly. It can be removed from the atmosphere when absorbed by plants. Removing carbon dioxide from the atmosphere reduces the greenhouse effect and global warming.

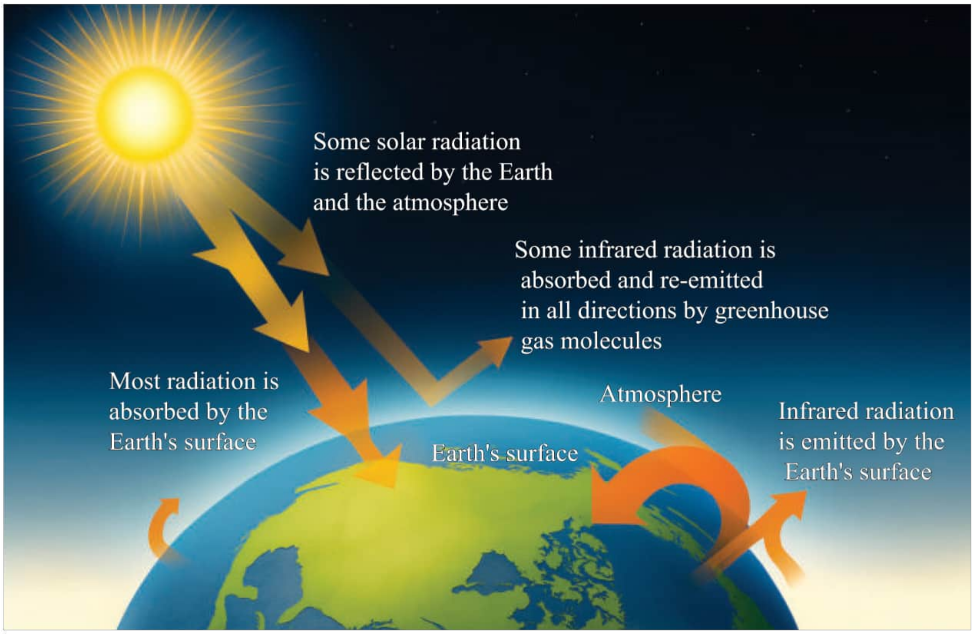


Fig.11.6: Greenhouse effect

ACITIVITY

Greenhouse Gas Emissions Investigation

Objectives:

To examine greenhouse gas emission sources, impacts, and personal habitsinfluencing emission levels.

Materials:

Chart paper and markers, internet access for research, drawing/ coloringmaterials, and worksheets for each student.

Procedure:

Begin by discussing the major greenhouse gases, including carbon dioxide (CO_2), methane (CH_4), and nitrous oxide (N_2O). Divide students into groups, assigning each aspecific greenhouse gas to research sources and impacts. Groups create charts summarising findings. Discuss the collective effect of greenhouse gases on global warming and climate change. Distribute worksheets listing daily activities, prompting students to identify emissions-contributing behaviours. Facilitate a discussion onpersonal habits for emission reduction. Encourage students to create individual actionplans to reduce their carbon footprint. Emphasise the importance of collective efforts inmitigating climate change.

Results:

Students explored the sources and impacts of greenhouse gases, as well as their own habits, through group research. They developed action plans that demonstrated increased awareness and commitment to reducing emissions

Conclusion:

The investigation raised awareness and prompted the adoption of proactive measures to reduce emissions. Engaging students cultivates collective efforts in combating climate change.

11.5 The Role of Sulphur, NO and NO_2 in the Formation of Acid Rain and Its Environmental Impact

Acid rain mainly forms when sulphur dioxide (SO_2) and nitrogen oxides (NO and NO_2) are released into the atmosphere from sources like power stations, vehicles, and industrial processes. These gases undergo chemical changes in the presence of oxygen and water vapour, producing strong acids that return to the Earth's surface as acidic precipitation, which includes rain, snow, fog, dew, and even dry particles.

Formation of Sulphuric Acid (H_2SO_4)

Sulphur dioxide (SO_2) reacts with oxygen and water in the atmosphere to form sulphuric acid:

$$SO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2SO_{3(aq)}$$

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$

Formation of Nitric Acid (HNO_3)

Nitric oxide (NO) is produced during high-temperature combustion. In the atmosphere, it oxidises to nitrogen dioxide (NO_2):

$$2NO + O_2 \longrightarrow 2NO_2$$

Nitrogen dioxide further reacts with water and oxygen to form nitric acid:

$$4NO_2 + 2H_2O + O_2 \longrightarrow 4HNO_3$$

Catalytic Role of Nitrogen Oxides

Nitrogen oxides not only produce nitric acid but also catalyse the oxidation of sulphur dioxide, thereby increasing acid rain formation. NO_2 oxidises SO_2 to SO_3, which then reacts with water to form sulphuric acid.

$$SO_2 + NO_2 \longrightarrow SO_3 + NO$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

The produced nitric oxide (NO) is re-oxidised to NO_2 by atmospheric oxygen, enabling the catalytic cycle to continue.

$$2NO + O_2 \longrightarrow 2NO_2$$

11.5.1 Environmental Impacts of Acid Rain

The presence of sulphuric acid in acid rain causes various harmful effects on the environment:

i) Acidification of Aquatic Ecosystems: It can increase the acidity of water bodies such as rivers, lakes, and oceans, which harms aquatic life. This may cause the death of fish, invertebrates, and other aquatic animals. It can also damage vegetation.

ii) Soil Degradation: Acid rain can leach essential nutrients from the soil, thereby decreasing fertility and negatively affecting plant growth.

iii) Vegetation Damage: Acid rain affects trees by causing leaf loss, damaging bark, and slowing growth, which weakens forests and decreases biodiversity.

iv) Infrastructure Corrosion: Acid rain damages buildings, monuments, and stone sculptures. It also erodes steel structures such as bridges and towers. This leads to increased maintenance costs and structural degradation, particularly for buildings constructed with limestone, marble, and concrete.

v) Human Health: Acid rain adversely affects human health by producing airborne particles that can lead to respiratory issues in both humans and animals. Children and those with pre-existing respiratory conditions, such as asthma, are especially at risk.

Self-Assessment

Explain how burning fossil fuels leads to both global warming and acid rain.

11.6 Strategies to Reduce the Effects of Major Environmental Issues

Some strategies to reduce the effects of major environmental issues:

11.6.1 Climate change

i) Planting Trees: Deforestation destroys approximately 2,000 trees every minute worldwide for agricultural and construction purposes, leading to an increase in carbon dioxide (CO_2) emissions. Dead trees no longer absorb CO_2, and their decay or combustion releases stored carbon as CO_2. Planting more trees is vital for reducing greenhouse gas emissions. During photosynthesis, trees not only absorb CO_2 but also produce essential compounds necessary for their growth.

Photosynthesis is a process in which carbon dioxide and water react in the presence of chlorophyll and light energy to produce glucose and oxygen. The balanced chemical equation for photosynthesis is:

$$6CO_2 + 6H_2O \xrightarrow[\text{Light Energy}]{\text{Chlorophyll}} C_6H_{12}O_6 + 6O_2$$

This process helps reduce greenhouse gas emissions and mitigate climate change by enhancing the natural absorption and storage of CO_2.

QUICK CHALLENGE

If deforestation increases, how does it impact atmospheric CO_2 levels and global temperature?

ii) **Reduction in Livestock Farming:** Livestock farming, especially cattle, releases methane during digestion and manure decomposition of manure. Reduction of methane emissions can be done by installing biogas plants.

iii) **Decreasing Use of Fossil Fuels:** Fossil fuels, including coal, oil, and natural gas, are major sources of greenhouse gas emissions. Transitioning from fossil fuels to renewable energy sources, such as solar, wind, geothermal, and hydropower, is vital for reducing overall greenhouse gas emissions.

iv) **Increasing Use of Hydrogen:** Hydrogen is regarded as a clean-burning fuel that produces only water vapour as a byproduct and does not directly emit greenhouse gases. Using hydrogen fuel cells in vehicles and industries can provide a cleaner alternative to fossil fuels.

v) **Renewable Energy:** Wind and solar energy are renewable sources that generate electricity without producing greenhouse gas emissions.

11.6.2 Acid Rain

a) Use of Catalytic Converters in Vehicles

In the context of acid rain, vehicle catalytic converters play a crucial role in reducing emissions of harmful compounds, including sulphur dioxide. These devices promote the complete combustion of fuel, resulting in lower overall emissions and a decrease in the formation of sulphuric acid in the atmosphere.

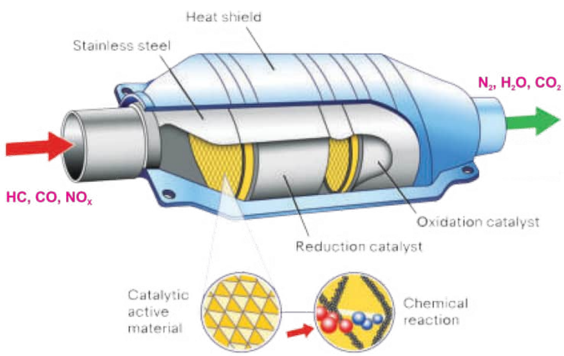


Fig. 11.7: Catalytic converter

An efficient catalytic converter (three-way catalytic converter) performs three functions:

(i) It oxidises CO to CO_2;

(ii) It reduces NO and NO_2 to N_2 and O_2;

(iii) It oxidises unburnt hydrocarbons to carbon dioxide and water.

$$2CO_{(g)} + O_{2(g)} \longrightarrow 2CO_{2(g)}$$

$$2NO_{(g)} \longrightarrow N_{2(g)} + O_{2(g)}$$

$$2CO_{(g)} + 2NO_{(g)} \longrightarrow 2CO_{2(g)} + N_{2(g)}$$

$$VOCs + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$$

b) Flue Gas Desulphurisation with Calcium Oxide

Another important technology, flue gas desulphurisation (FGD) using calcium oxide, is used in power stations to eliminate sulphur dioxide from flue gases before their release, significantly reducing sulphur emissions from large-scale combustion processes.

Real-World Connections

Flue gas desulphurisation is used in power plants to trap sulphur dioxide before it escapes into the air, a key step in reducing acid rain.

11.7 Environmental Risks and Long-Term Respiratory Health

High-risk situations that may result in respiratory issues due to prolonged exposure to pollutants include:

i) **Indoor Pollutants:** Radon gas, tobacco smoke, mould spores, and household chemicals can lead to respiratory issues such as asthma, allergies, and, in severe cases, lung cancer when present indoors.

ii) **Outdoor Air Pollution:** Pollution from vehicles, industries, and other sources can lead to asthma, bronchitis, reduced lung function, and an increased risk of lung cancer.

iii) **Smoking:** Smoking is a significant risk factor for chronic respiratory diseases. Tobacco smoke contains harmful chemicals that can lead to lung diseases, such as lung cancer and chronic bronchitis, when inhaled regularly.

iv) **Secondhand Smoke:** Non-smokers exposed to tobacco smoke face a higher risk of developing asthma, bronchitis, and lung cancer.

v) **Allergens:** Extended exposure to allergens such as pollen, dust mites, and animal dander can result in lung diseases that may lead to allergic rhinitis and asthma.

vi) **Climate Change:** The effects of climate change, such as poor air quality, higher pollen levels, and extreme weather conditions, can further affect respiratory health and overall well-being over time.

Real-World Connections

Reducing sulphur emissions not only protects monuments like the Badshahi Mosque in Lahore from corrosion but also saves millions of rupees in annual repair costs.

11.8 Test for the Presence of Water Using Anhydrous Copper(II) Sulphate

This test is frequently used in chemistry laboratories to identify the presence of water.

Introduction

Anhydrous copper(II) sulphate is a white compound that reacts with water. This reaction is known as hydration, where anhydrous copper sulphate absorbs water molecules and changes into hydrated copper sulphate. Anhydrous means "without water".

$$\underset{\text{White}}{CuSO_{4(s)}} + 5H_2O_{(l)} \longrightarrow \underset{\text{Blue}}{CuSO_4.5H_2O_{(s)}}$$

Materials

i) Anhydrous copper sulfate crystals
ii) The substance being tested
iii) Test tube
iv) Dropper

Fig. 11.8: Test for the presence of water

Procedure

I) Place a small amount of anhydrous copper sulphate into a test tube.
ii) Add a few drops of the substance under examination.
iii) Gently shake the test tube to mix its contents.
iv) If the crystals turn blue, it indicates that water is present in the tested substance.
v) If the crystals stay white, the tested substance has low or no water content.

Table 11.2: Water distribution on earth

Water Type	Percentage of Total Water	Explanation
Seas and oceans	97.2%	Saltwater covering 71% of Earth's surface
Polar Ice Caps and Glaciers	2.15%	Frozen freshwater reserves
Fresh Liquid Water	0.63%	Lakes, rivers, and groundwater
Water Vapour	0.001%	Water in the atmosphere

Note The test is not specific to water. Other substances, such as alcohols, can also cause the anhydrous copper(II) sulphate to turn blue.

11.9 Testing Water Purity Using Melting and Boiling Points

Both melting and boiling points can be used to measure the purity of water. Pure water has specific melting and boiling points under standard atmospheric pressure. The melting point of pure water is 0°C (32°F), while its boiling point is 100°C (212°F). Any significant deviation from the standard boiling point may indicate the presence of impurities that affect the melting and boiling points. Typically, impurities reduce the melting point and increase the boiling point of the water.

Do You Know?

Although 71% of the Earth's surface is covered with water, less than 1% is readily accessible freshwater fit for drinking.

Melting Point Test

i) Pour a sample of water into a clean beaker.

ii) Slowly cool the water below its freezing point.

iii) Note the temperature at which water solidifies.

Result: Water that freezes at 0°C (32°F) is considered pure. Any deviation from 0°C indicates the presence of impurities in the water.

Boiling Point Test

i) Place a sample of water into a clean beaker.

ii) Heat the water until it boils.

iii) Observe the temperature at which water boils.

Result: Water that boils at 100°C (212°F) is considered pure. A deviation from 100°C indicates the presence of impurities in the water.

FUN FACT

One litre of water can produce over 1,200 litres of steam when boiled, illustrating the significant difference in the spacing of gas molecules compared to liquids.

Table 11.3: Comparison of tap water and distilled water

Feature	Tap Water	Distilled Water
Definition	Water supplied through municipal systems is treated for safe domestic use but may contain minerals and trace impurities.	Water purified through distillation is free from minerals, impurities, and microorganisms.
Impurities	Varies (minerals, chemicals, microorganisms)	Virtually none
Taste	Varies depending on minerals and chemicals	Neutral
Minerals	May contain beneficial minerals	Lacks minerals
Suitability for drinking	Generally safe for drinking (depending on quality)	Not recommended for long term - consumption
Common uses	Drinking, cooking, cleaning, and bathing	Laboratory experiments, steam irons, batteries, and cleaning

11.10 Water from Natural Sources

Water from natural sources can contain both useful and harmful substances. Some examples of these substances include:

11.10.1 Beneficial Substances

i) Dissolved oxygen: It is essential for the survival of aquatic life and support the growth of some microorganisms.

ii) Minerals: It is important for human health, playing key roles in various physiological functions. Some examples of minerals include: calcium, sodium, potassium, iron, and magnesium.

11.10.2 Harmful Substances

i) Metal Compound: It is toxic when present in high concentrations. It may adversely disrupts ecosystem.

ii) Plastics: It can harm wildlife and pollute the environment, leading to ecological imbalances and threats to marine life.

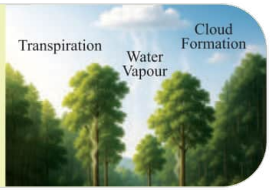
iii) Sewage: It is the mixture of wastes from the human body and used water. It contains harmful bacteria, viruses and organic matter. It poses the risk of waterborne diseases.

iv) Harmful Microbes: They are the microorganism such as bacteria and viruses and can cause disease. They can cause waterborne disease and pose health risks to humans and animals, especially when water is contaminated with pathogens.

v) Nitrates and phosphates from fertilisers and detergents: The runoff from agricultural and wastewater contains fertilisers and detergents. They can contribute to algae blooms and water pollution.

FUN FACT

A single large tropical tree releases up to 400 litres of water vapour into the air each day through transpiration. This process helps form clouds and cause rainfall. Therefore, forests are vital in maintaining Earth's water cycle.



11.11 Treatment of Domestic Water

The method in which domestic water is made fit for drinking and other household purposes by removing impurities is known as water treatment. It provides safe, clean, and pleasant-tasting water for our daily needs. Water treatment is a multi-step process.

i) Pretreatment: Water is passed through screens to remove large debris such as leaves, sticks, etc.

ii) Coagulation: Chemicals such as aluminum sulphate or potash alum is added to which causes small suspended particle impurities to clump together. This makes colloidal clay precipitates as a solid so that it can be filtered off.

iii) Sedimentation: The water is allowed to rest in sedimentation tanks, allowing the large or heavy particles to settle out.

vi) Filtration: Water is filtered through layers of sand and gravels to remove any fine particles and microorganisms.

v) Carbon Adsorption: Sometimes water is passed over beds of activated charcoal to remove unpleasant tastes, odors, and certain organic contaminants.

vi) Disinfection: Chlorine or compounds of chlorine is added to kill bacteria, viruses and other harmful microorganisms.

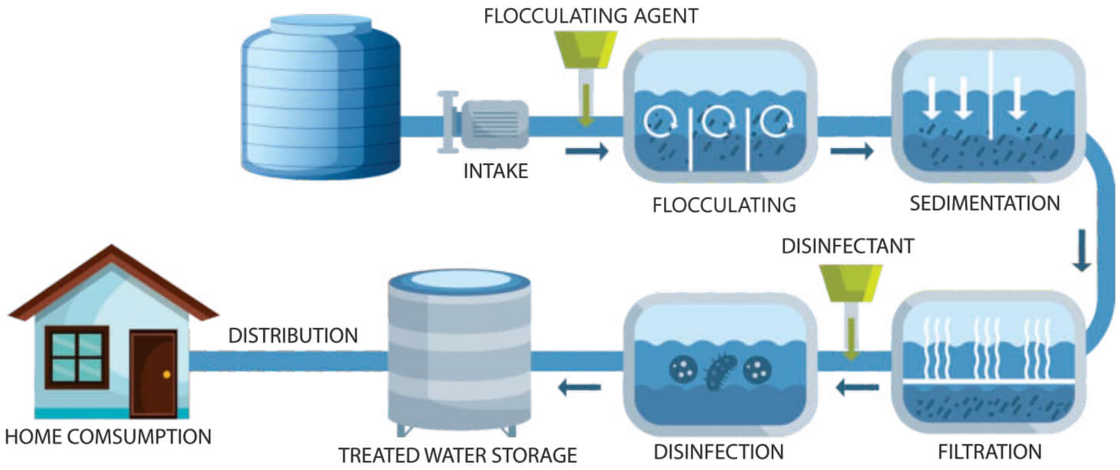


Fig. 11.9: Treatment of water

11.12 Waterborne Diseases

Waterborne diseases are illnesses caused by consuming or coming into contact with water contaminated by various harmful substances. These contaminants can have serious consequences for living organisms, including humans, and can lead to a range of health issues.

11.12.1 Common Waterborne Diseases

i) Gastrointestinal Diseases (Cholera and Diarrhoea): Intestinal diseases like cholera and diarrhoea are the most common waterborne illnesses, killing thousands of people each year.

ii) Hepatitis: Hepatitis is a disease that affects the liver and is primarily caused by a virus, often transmitted through contaminated water.

iii) Gastrointestinal Tract (GIT) Worms: Worms such as roundworms, hookworms, tapeworms, pinworms, and whipworms can be transmitted through contaminated water, resulting in health issues such as diarrhoea, anaemia, itching, and irritation of the gastrointestinal tract and rectum.

iv) Typhoid Fever: A bacterial disease spread through contaminated water sources, posing a significant health rf not treated properly.

11.13 Water Pollution: Negative Effects and Prevention Methods

Water pollutants can harm aquatic life, ecosystems, and human health. Below are some key negative impacts of water pollutants and methods to prevent them:

11.13.1 Adverse Effects of Water Pollutants on Life

i) Human Health Impacts: Contaminated water can spread diseases and lead to poisoning from chemical pollutants, including pesticides, mercury, and arsenic. It can also disrupt hormone function, contributing to developmental abnormalities and reproductive issues.

ii) Environmental Impacts: Water pollutants can harm aquatic organisms, and disrupt ecosystems, food chains, and biodiversity. Excessive nutrients can cause algal blooms and create "dead zones" devoid of aquatic life. Toxic pollutants can accumulate in fish and shellfish, posing risks to human health.

11.13.2 Ways to Avoid the Negative Effects of Water Pollutants

There are various ways to mitigate the adverse effects of water pollutants, such as implementing effective water treatment methods methode, promoting good hygiene and sanitation practices, enforcing regulations on industrial waste disposal, educating the public about the dabgers of contaminated water, investing in water infrastructure, and practising conservation efforts to protect water sources and ecosystems.

Implementing these measures, can help reduce the harmful effects of water pollution on human health and the environment.

11.14 Water Scarcity

Water scarcity refers insufficient freshwater necessary to meet the needs of both people and ecosystems in a specific area. Although water covers around 71% of the Earth's surface, most of it is unsuitable for human consumption. Freshwater, the only source of potable water, exists in very limited quantities. This scarcity is worsened by human activities and inadequate planning, resulting in waste and unsustainable usage patterns.

Fig. 11.10: Water scarcity

11.14.1 Importance of Water

Water is a vital resource used in nearly every human activity, including drinking, washing, cooking, and cleaning. According to the United Nations, a person needs a minimum of 50 liters of water per day to meet basic needs for hygiene, cooking, and drinking. However, many people worldwide do not receive this amount, resulting in severe health and social consequences. Insufficient water intake can result in kidney problems, constipation, and various mental and physical health issues. Therefore, preserving usable water is crucial for a healthy and sustainable life.

11.14.2 Water Scarcity in Pakistan

Pakistan is facing a severe water crisis, ranking 14th among the 17 "extremely high-water risk" countries. Over 80% of Pakistan's population suffers from severe water scarcity for at least one month each year. Moreover, the country's groundwater resources are being severely overdrawn, primarily for irrigation purposes.

11.14.3 Causes of Water Scarcity in Pakistan

i) Rapid Population Growth and Urbanisation: The population has risen significantly, placing tremendous pressure on water resources.

Do You Know?

Over 80% of Pakistan's freshwater is used in agriculture, yet inefficient irrigation methods waste nearly half of it before reaching the crops.

ii) Climate Change: Pakistan is highly vulnerable to climate change, experiencing altered monsoon patterns, receding glaciers, rising temperatures, and frequent floods and droughts.

iii) Poor Agricultural Practices: Four major crops (rice, wheat, sugarcane, and cotton) consume over 80% of water resources, yet they contribute only 5% to GDP. Inefficient irrigation methods lead to significant water wastage.

iv) Inefficient Infrastructure: The ageing and poorly maintained irrigation systems result in significant water losses.

v) Water Pollution: Industrial, agricultural, and domestic waste pollute freshwater sources, leading to health problems and reducing water availability.

vi) Inter-Provincial Water Distribution Issues: Disputes regarding water distribution among provinces worsen the water crisis.

11.14.4 Solutions to Water Scarcity in Pakistan

i) Political Ownership and Governance: The entire political leadership and stakeholders must take ownership of the water crisis by implementing effective policies and governance practices to address it.

ii) Improved Water Management: Invest in upgrading and maintaining water infrastructure, implement efficient water metering and pricing mechanisms, and enhance institutional capacity for integrated water resources management.

iii) Water Conservation and Efficiency: Promote water-efficient agricultural practices, adopt water-saving technologies, and enhance public awareness of water conservation.

iv) Diversifying Water Sources: Explore alternative sources such as groundwater recharge, rainwater harvesting, and wastewater treatment and reuse. Invest in desalination plants to make use of coastal resources.

v) Recycling Wastewater: Following the examples of Israel and Singapore, we should recycle wastewater for irrigation and industrial uses.

vi) Climate Change Adaptation: Implement climate-smart agricultural practices, improve early warning systems, and promote nature-based solutions to regulate the hydrological cycle.

vii) Water Policies and Reforms: Implement comprehensive water policies at both national and provincial levels, addressing water quality issues, establishing clear targets, and ensuring gender inclusion.

viii) Addressing Inter-Provincial Disputes: Develop fair water distribution mechanisms and enhance collaboration among provinces.

11.15 Fertilisers

Fertilisers are substances that provide essential nutrients to plants, promoting their growth and development. They increase crop production by supplementing the soil with vital elements necessary for healthy plant life. Urea, ammonium salts, and nitrates are common examples of fertilisers, each offering different levels and forms of nitrogen. The best type of fertiliser for your specific needs will depend on several factors, including the climate, your budget, crop type, and soil conditions. Pakistan is an agricultural country that requires fertilisers to enhance crop production.

11.15.1 Urea

Urea is a high-quality nitrogenous fertiliser. It is a white crystalline solid and contains 46% nitrogen. It is the most concentrated source of nitrogen among commonly used fertilisers. Urea is readily soluble in water and is absorbed by plants through their roots. This is the most widely used fertiliser in Pakistan.

Fig. 11.11: Urea fertiliser

FUN FACT

The white crystals of urea were the first organic compound synthesised in a lab by Friedrich Wohler in 1828, demonstrating that organic substances can be produced artificially.

11.15.2 Ammonium Salts

Ammonium salts, such as ammonium nitrate and ammonium sulphate, are significant nitrogen sources for plants, typically containing 21-33% nitrogen. Although ammonium salts are generally less soluble, they are still readily absorbed by plants and are used for many crops, except for paddy rice.

Fig. 11.12: Ammonium chloride fertiliser

Real-Life Application

Nitrogenous fertilisers such as urea and ammonium nitrate help increase grain yields, but must be applied carefully to prevent groundwater pollution.

11.15.3 Nitrates

Nitrates, including calcium nitrate ($Ca(NO_3)_2$) and potassium nitrate (KNO_3), are vital plant nutrients. They are highly soluble in water, making them readily available for absorption by plant roots. Nitrates typically contain 13-16% nitrogen, which is essential for plant growth and development. However, nitrates are also more prone to leaching, which raises environmental concerns. Leaching occurs when water washes away nitrates and other soluble nutrients from

Fig. 11.13: Calcium nitrate fertiliser

the plant root zone, where they can contaminate groundwater and surface water. Nitrate leaching can pollute groundwater, posing risks to both human health and ecosystems.

KEY POINTS

1. The atmosphere is a vital protective layer surrounding the Earth, sustaining life by regulating temperatures, blocking harmful solar radiation, and facilitating essential processes such as photosynthesis.
2. The major components of the atmosphere are nitrogen (78.08%) and oxygen (20.94%), with minor components such as argon, carbon dioxide, and trace gases playing vital roles in various ecological processes.
3. The atmosphere is divided into four main layers: the troposphere, stratosphere, mesosphere, and thermosphere, each with distinct altitude ranges and temperature characteristics.
4. Major sources of air pollutants include carbon dioxide, carbon monoxide, methane, oxides of nitrogen, oxides of sulfur, and ground-level ozone, all of which have significant consequences for the environment and human health.
5. Carbon dioxide (CO_2) is non-toxic, but excessive amounts strengthen the greenhouse effect, causing global warming, melting ice caps, rising sea levels, and severe climatic changes.
6. Carbon monoxide (CO) is a colourless and odourless gas produced by the incomplete combustion of fuels. It is extremely toxic as it inhibits oxygen transport in the blood, posing serious health and environmental risks.
7. Methane (CH_4) is a potent greenhouse gas emitted from agriculture, wetlands, and fossil fuel industries. It significantly contributes to global warming and climate change.
8. Oxides of nitrogen (NO_x), primarily nitric oxide (NO) and nitrogen dioxide (NO_2), are generated from vehicle emissions and industrial combustion. They cause respiratory issues, contribute to smog formation, and result in acid rain.
9. Oxides of sulphur (SO_x), particularly sulphur dioxide (SO_2) and sulphur trioxide (SO_3), are released from burning sulphur-rich fuels. They form acid rain that damages buildings, vegetation, aquatic ecosystems, and human health.
10. Ozone (O_3) is beneficial in the stratosphere, but ground-level ozone contributes to photochemical smog, leading to respiratory problems, eye irritation, and damage to vegetation and materials.
11. Water is essential for all living organisms, playing a vital role in many biological processes, such as digestion, joint lubrication, temperature regulation, and nutrient transportation.
12. Although the Earth has plenty of water, only a small part (0.63%) is freshwater suitable for drinking. Most water is in seas and oceans (97.2%), polar ice caps and glaciers (2.15%), and water vapour in the atmosphere (0.001%).
13. Glaciers and ice caps contain 69% of the world's freshwater, while groundwater makes up 30.1%, and surface water (rivers, lakes, reservoirs) accounts for approximately 0.9%.
14. A common laboratory test to detect water is the anhydrous copper(II) sulphate test. In this test, anhydrous copper sulphate (white) turn blue when hydrated by water.

15. The purity of water can also be checked by observing its melting and boiling points. Pure water freezes at 0°C and boils at 100°C; any deviation from these values indicates the presence of impurities.
16. Tap water is sourced from natural supplies, often contains minerals, and is suitable for various domestic uses. Distilled water, on the other hand, is produced by distillation, lacks minerals, and is primarily used in laboratory experiments and appliances such as steam irons.
17. Water treatment involves several stages, including pretreatment, flocculation, sedimentation, filtration, carbon adsorption, and disinfection, to ensure the production of safe and clean water for household use.
18. Contaminated water can cause waterborne diseases such as cholera, hepatitis, gastrointestinal disorders, and typhoid fever, emphasising the importance of clean water for public health.
19. In Pakistan, water scarcity has become a serious issue due to factors such as rapid population growth, climate change, inefficient water management, water pollution, and limited water storage capacity.
20. In agriculture, fertilisers such as urea, ammonium salts, and nitrates are frequently used to improve crop yield by supplying vital nutrients like nitrogen to plants.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The primary function of the atmosphere is:
   (a) absorbing cosmic rays (b) generating heat
   (c) sustaining life on Earth (d) producing greenhouse gases

ii) The gas considered a major contributor to global warming is:
   (a) nitrous oxide (b) carbon monoxide
   (c) carbon dioxide (d) sulphur dioxide

iii) The main source of methane emissions is:
   (a) Wetlands and rice paddies (b) Agriculture and livestock farming
   (c) Volcanoes and forest fires (d) Industrial waste gases

iv) The main gas responsible for trapping infrared radiation and contributing to global warming is:
   (a) methane (b) carbon dioxide
   (c) nitrous oxide (d) ozone

v) The main product formed when sulphur dioxide reacts with water vapour in the atmosphere is:
   (a) sulphuric acid (b) nitric acid
   (c) carbonic acid (d) hydrochloric acid

vi) The poisonous gas produced by the incomplete combustion of carbon-based fuels is:
(a) carbon monoxide (b) sulphur dioxide
(c) methane (d) nitrogen dioxide

vii) The major component of Earth's atmosphere is:
(a) oxygen (b) nitrogen
(c) carbon Dioxide (d) argon

viii) The percentage of Earth's water found in seas and oceans is:
(a) 71% (b) 97.2%
(c) 2.15% (d) 0.63%

ix) The test commonly used in chemistry laboratories to detect the presence of water is:
(a) boiling Point Test
(b) melting Point Test
(c) anhydrous Copper(II) Sulphate Test
(d) coagulation Test

x) The harmful substance associated with water pollution from fertilisers and detergents is:
(a) dissolved oxygen (b) harmful microbes
(c) nitrates and phosphates (d) beneficial mineral

B. Restricted Response Questions (RRQs)

i) Write the balanced chemical equation for the formation of carbon dioxide from the combustion of methane.

ii) Which oxide of nitrogen is reddish-brown and has a pungent smell?

iii) List four harmful effects of acid rain on the environment.

iv) What chemical process leads to the formation of ground-level ozone?

v) Describe the role of sulphur dioxide in the formation of acid rain.

vi) Why is distilled water not recommended for long-term human consumption?

vii) Why is nitrate leaching a concern for the environment?

viii) Which sector consumes over 80% of Pakistan's freshwater resources?

C. Extended Response Questions (ERQs)

i) Describe the composition of clean, dry air and explain the significance of two of its major components.

ii) Explain the process of global warming and identify its major causes.

iii) Identify and explain any four major pollutants found in the atmosphere, and state one effect of each.

iv) Explain the significance of water for living organisms and why humans need to consume water daily.

v) Outline the key stages involved in the treatment of domestic water and explain the purpose of each stage.

vi) Identify and explain the major causes of water scarcity in Pakistan.

UNIT 12

ORGANIC CHEMISTRY

Student Learning Outcomes (SLOs)

After studying this chapter, the students will be able to:

E-01 Define organic compounds with examples.

E-02 Describe organic molecules as either straight-chained, branched or cyclic.

E-03 Explain why a systematic method of naming chemical compounds is necessary.

E-04 State that a structural formula is an unambiguous description of the way the atoms in a molecule are arranged, including $CH_2=CH_2$, CH_3CH_2OH, CH_3COOCH_3.

E-05 Identify and draw structural formulae for molecules.

E-06 Interpret general formulae of compounds in the same homologous series including alkanes, alkenes, alkynes, alcohols and carboxylic acids.

E-07 Define structural isomers as compounds with the same molecular formula, but different structural formulae, including C_4H_{10} as $CH_3CH_2CH_2CH_3$ and $CH_3CH(CH_3)CH_3$ and C_4H_8 as $CH_3CH_2CH=CH_2$ and $CH_3CH=CHCH_3$.

E-08 Identify a functional group as an atom or group of atoms that determine the chemical properties of a homologous series including that for alkyl halides alcohols, aldehydes, ketones, phenols, carboxylic acids, amine, esters, and amide.

E-09 Describe the general characteristics of a homologous series . (These can include:
- (a) having the same functional group
- (b) having the same general formula
- (c) differing from one member to the next by a $-CH_2-$ unit
- (d) displaying a trend in physical properties
- (e) sharing similar chemical properties.)

E-10 State that a saturated compound has molecules in which all carbon–carbon bonds are single bonds.

E-11 State that an unsaturated compound has molecules in which one or more carbon– carbon bonds are not single bonds.

E-12 State that the bonding in alkanes is single covalent and that alkanes are saturated hydrocarbons.

E-13 Describe the properties of alkanes as being generally unreactive, except in terms of combustion and substitution by chlorine.

E-14 State that in a substitution reaction one atom or group of atoms is replaced by another atom or group of atoms.

E-15 Describe the substitution reaction of alkanes with chlorine as a photochemical reaction, and draw the structural or displayed formulae of the products, limited to monosubstitution.

E-16 Describe, using symbol equations, preparation of alkanes from cracking of larger hydrocarbons, hydrogenation of alkenes and alkynes, and reduction of alkyl halides.

Introduction

Organic chemistry is the branch of chemistry that studies carbon-based compounds, concentrating on their structure, properties, and reactions. It explains essential life processes such as photosynthesis, respiration, and the synthesis of proteins, lipids, and nucleic acids. Numerous medicines, including antibiotics, analgesics, anaesthetics, and antivirals, are organic compounds used to treat various diseases.

Organic chemistry is also essential in developing materials such as plastics, rubbers, fibres, and composites, as well as in producing fuels, dyes, agrochemicals, and fragrances. Organic compounds, including fossil and biofuels, are key energy sources.

Hydrocarbons, composed only of carbon and hydrogen, are the simplest organic compounds. They are the main constituents of fuels such as petrol, diesel, and natural gas, and act as raw materials in the production of plastics, pharmaceuticals, synthetic fibres, and lubricants. Their usefulness makes them essential to modern industry and technology.

12.1 Organic Compounds

Organic compounds are chemical substances that are chemical substances that contain carbon and many other elements, such as hydrogen, oxygen, nitrogen, sulphur, phosphorus, and halogens. A few carbon-containing compounds that are not classified as organic include carbonates, bicarbonates, carbides, cyanides, cyanates, CO, CO_2, and CS_2. Carbon is the primary element in organic compounds and is essential to life. All living organisms are made up of carbon compounds. Some examples of organic molecules are methane, acetylene, benzene, carbohydrates, proteins, lipids, and nucleic acids.

MISCONCEPTION Vs REALITY

Myth: Only living organisms can produce organic compounds.

Reality: Chemists can synthesise organic compounds in laboratories.

12.1.1 Classification of Organic Compounds/Molecules

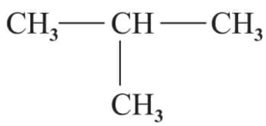
Based on their structural arrangement, organic molecules can be classified into three main types: straight-chained, branched, or cyclic.

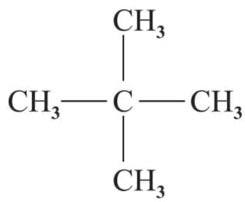
i) Straight-chained hydrocarbons show a linear, unbranched arrangement of carbon atoms, resulting in a continuous chain. Examples of straight-chained molecules are:

$CH_3—CH_2—CH_2—CH_3$
n-Butane

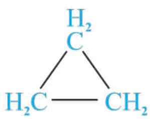
$CH_3—CH_2—CH_2—CH_2—CH_3$
n-Pentane

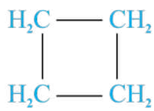
ii) Branched hydrocarbons contain a main carbon chain with one or more branches extending from it. For example:

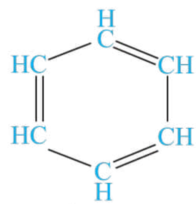

Isobutane


Neopentane

iii) Cyclic or closed chain hydrocarbons consist of a closed ring of carbon atoms. For example:


Cyclopropane


Cyclobutane


Benzene

In cyclic hydrocarbons, each carbon atom in the ring is bonded to at least two other atoms. These hydrocarbons range from simple three-membered rings (such as cyclopropane) to larger, more complex structures.

12.2 Saturated and Unsaturated Hydrocarbons

Hydrocarbons are the most basic organic compounds made up of carbon and hydrogen atoms only These carbon and hydrogen atoms are connected by covalent bonds. Examples include methane (CH_4), ethylene (C_2H_4), acetylene (C_2H_2), benzene (C_6H_6), and so on.

FUN FACT
The pleasant smell of petrol comes from aromatic hydrocarbons such as benzene derivatives.

Hydrocarbons can be classified into two main types based on the presence of bonds: saturated and unsaturated.

12.2.1 Saturated Hydrocarbons

Hydrocarbons that contain only single bonds between carbon atoms are called saturated hydrocarbons. They are also known as alkanes. Alkanes represent the simplest family of organic compounds. Petroleum and natural gas are typically composed of alkanes. The simplest alkane is methane, which is the primary component of natural gas. The term "saturated" refers to the fact that the four valencies of the carbon atoms in these molecules are fully satisfied, meaning that no further atoms or groups of atoms can be added to the carbon skeleton.

The names of the first four straight-chain alkanes are listed in the table below:

Table 12.1: The first four straight-chain alkanes

IUPAC Name	Number of C Atoms	Molecular Formula	Structural Formula
Methane	1	CH_4	CH_4
Ethane	2	C_2H_6	$CH_3 - CH_3$
Propane	3	C_3H_8	$CH_3 - CH_2 - CH_3$
Butane	4	C_4H_{10}	$CH_3 - CH_2 - CH_2 - CH_3$

Their general formula is C_nH_{2n+2}, where "n" is the number of carbon atoms in the molecule. Their names all end with the suffix -ane.

12.2.2 Unsaturated hydrocarbons

Hydrocarbons that contain at least one double or triple bond between two carbon atoms are termed unsaturated hydrocarbons. In these compounds, the four valencies of carbon are not fully satisfied, allowing for the addition of other atoms or groups of atoms to carbon atoms. These hydrocarbons include alkenes and alkynes.

i). Alkenes: Alkenes are unsaturated hydrocarbons that contain one or more carbon-carbon double bonds (C=C) in their structures. They are found in small amounts in natural gas, coal gas, and petrol. Alkenes have a general formula of C_nH_{2n}, where **n** represents the number of carbon atoms. Ethylene is the first member of the series, and some of the simplest alkenes are listed in the table 12.2.

Table 12.2: Structural formulae of alkenes

Alkenes	Number of C Atoms	Molecular Formula	Structural Formula
Ethene	2	C_2H_4	$CH_2 = CH_2$
Propene	3	C_3H_6	$CH_2 = CH - CH_3$
1-Butene	4	C_4H_8	$CH_2 = CH - CH_2 - CH_3$
2-Butene	4	C_4H_8	$CH_2 - CH = CH_2 - CH_3$

Interesting Information

During the ripening of bananas, ethylene gas is produced, which accelerates the ripening process of the fruit and can lead to faster spoilage. Ethylene gas accelerates and enhances the decay process of nearby fruits. It is, therefore, recommended that bananas should not be stored alongside other fruits.

ii) Alkynes: Alkynes are unsaturated hydrocarbons that contain one or more carbon-carbon triple bonds in their structures. They have a general formula of C_nH_{2n-2}, where n represents the number of carbon atoms in the molecule. Alkynes are less common in nature than alkenes, although more than a thousand different alkynes have been isolated from natural sources.

Table 12.3: Structural formulae of alkynes

Alkynes	Number of C Atoms	Molecular Formula	Structural Formula
Ethyne	2	C_2H_2	$CH \equiv CH$
Propyne	3	C_3H_4	$CH \equiv C - CH_3$
1-Butyne	4	C_4H_6	$CH \equiv C - CH_2 - CH_3$
2-Butyne	4	C_4H_6	$CH_3 - C \equiv C - CH_3$

12.3 Structural Formula

A structural formula is a two-dimensional representation of a three-dimensional molecule that clearly illustrates the arrangement and connectivity of atoms and chemical bonds within the molecule.

Table 12.4: Structural formulae of molecules

Organic Compounds	Molecular Formula	Structural Formula
Ethene	C_2H_4	$CH_2 = CH_2$
Ethyl alcohol	C_2H_6O	$CH_3 - CH_2 - OH$
Methyl acetate	$C_3H_6O_2$	$CH_3 - \overset{\overset{\displaystyle O}{\Vert}}{C} - OCH_3$

12.4 Homologous Series

A group of organic compounds where each member differs from the next by a methylene group ($-CH_2-$) is known as a homologous series. The members of a homologous series are referred to as homologues. All members of a homologous series are represented by the same general formula. For example, the general formulae for alkanes, alkenes, alkynes, and alcohols are C_nH_{2n+2}, C_nH_{2n}, C_nH_{2n-2} and $C_nH_{2n+1}OH$, respectively. The integer "n" shows the number of carbon atoms in the molecules. For example, butane ($CH_3CH_2CH_2CH_3$) and pentane ($CH_3CH_2CH_2CH_2CH_3$) are homologues.

Table 12.5: The homologous series of alkanes

Name of Alkanes	Number of C Atoms	Molecular Formula	Condensed Structural Formula
Methane	1	CH_4	CH_4
Ethane	2	C_2H_6	CH_3CH_3
Propane	3	C_3H_8	$CH_3CH_2CH_3$
Butane	4	C_4H_{10}	$CH_3CH_2CH_2CH_3$
Pentane	5	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$
Hexane	6	C_6H_{14}	$CH_3CH_2CH_2CH_2CH_2CH_3$
Heptane	7	C_7H_{16}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$

12.4.1 General Characteristics of Homologous Series

QUICK CHALLENGE

How many hydrogen atoms are in an alkane with eight carbon atoms?

1. Every member of the homologous series has the same structural formula.
2. All members of the homologous series have the same functional groups.
3. All members of the homologous series share the same general formula.
4. They have similar chemical properties because of their identical functional groups.
5. They have distinct physical properties due to the different molecular masses of the members of the homologous series.
6. The molecular formula of each successive homologue differs by a "—CH_2—" group.
7. Every compound in the series contains the same type of elements.

CHEMISTRY IN LIFE

Carboxylic acids and alcohols play a crucial role in forming fats and oils in living organisms.

12.5 Functional Groups

A functional group is an atom or a specific group of atoms within an organic molecule that determines its chemical properties and reactivity. It is the reactive part of the compound, while the rest of the molecule, mainly the hydrocarbon chain, remains comparatively inert. Compounds with the same functional group show similar chemical properties and belong to the same homologous series. Double and triple bonds also act as functional groups that influence reactivity. If one functional group is replaced by another, the chemical behaviour and properties of the molecule change.

For example, the chemical properties of methanol (CH_3OH), ethanol (CH_3CH_2OH), and propanol ($CH_3CH_2CH_2OH$) are similar because they all contain the same functional group: the hydroxyl (−OH) group. Methanol (CH_3OH) and methanoic acid (HCOOH) have different chemical properties because they have different functional groups.

Importance

This concept is important in understanding the:

1. Classification and nomenclature of organic compounds.
2. The behaviour of organic compounds, including those in homologous series such as alcohols, aldehydes, ketones, phenols, carboxylic acids, amines, esters, and amides.

Interesting Information

The term "functional group" was first introduced in the 19th century when chemists observed that specific atoms or groups of atoms consistently produced similar chemical reactions, regardless of the carbon framework to which they were attached.

Table 12.6: Common functional groups and classes of organic compounds

Class of the Compound	Functional Group	Common Formula	Examples
Alkenes	$C{=}C$	$R{-}C{=}C{-}R'$	Ethene, $CH_2{=}CH_2$
Alkynes	$C{\equiv}C$	$R{-}C{\equiv}C{-}R'$	Ethyne, $CH{\equiv}CH$
Alcohols	$-OH$	$R{-}OH$	Ethanol, CH_3CH_2OH
Aldehydes	$-CHO$	$R{-}CHO$	Acetaldehyde, CH_3CHO
Ketones	$-CO-$	$R{-}CO{-}R'$	Acetone, CH_3COCH_3
Carboxylic Acids	$-COOH$	$R{-}COOH$	Acetic acid, CH_3COOH
Esters	$-COOR$	$R{-}COOR'$	Methyl acetate, CH_3COOCH_3
Ethers	$-O-$	$R{-}O{-}R'$	Dimethyl ether, CH_3OCH_3
Amines	$-NH_2$ (primary), $-NHR$ (secondary), $-NR_2$ (tertiary)	$R{-}NH_2$, R_2NH, R_3N	Methylamine, CH_3NH_2
Amides	$-CONH_2$	$R{-}CONH_2$	Acetamide, CH_3CONH_2
Halides (Alkyl halides)	$-X$ (F, Cl, Br, I)	$R{-}X$	Chloromethane, CH_3Cl
Nitriles	$-C{\equiv}N$	$R{-}C{\equiv}N$	Acetonitrile, CH_3CN

12.6 Interpretation of Homologous Series through General Formulae

The general formulae indicate the ratio of carbon to hydrogen atoms in the molecules and help in understanding the similarities and differences among compounds within the same homologous series. Some general formulae and their interpretation for the classes of these compounds are:

12.6.1 Alkanes

The general formula for alkanes is C_nH_{2n+2}, where "n" represents the number of carbon atoms in the molecule and "(2n+2)" indicates the number of hydrogen atoms. Alkanes have twice as many hydrogen atoms as carbon atoms. For example, butane (C_4H_{10}) has four carbon atoms (n = 4) and ten hydrogen atoms ($2 \times 4 + 2 = 10$).

12.6.2 Alcohols

The general formula for alcohols is $C_nH_{2n+1}OH$, where "n" denotes the number of carbon atoms in the molecule, and "(2n+1)" indicates the number of hydrogen atoms. Alcohols have one more hydrogen atom than carbon atoms, plus the hydroxyl (OH) group. For example, ethanol (CH_3CH_2OH) has two carbon atoms (n=2) and five hydrogen atoms (2×2+1=5), plus the OH group.

12.6.3 Carboxylic Acids

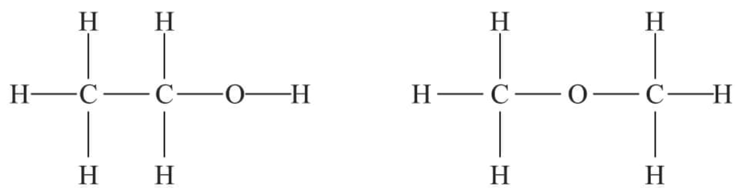
The general formula for carboxylic acids is $C_nH_{2n+1}COOH$, where "n" represents the number of carbon atoms in the molecule, and "2n+1" represents the number of hydrogen atoms. Carboxylic acids have one more hydrogen atom than twice the number of carbon atoms, plus the carboxyl (COOH) group. For example, ethanoic acid (CH_3COOH) has two carbon atoms (n=2) and five hydrogen atoms (2×2+1=5), plus the COOH group.

Self-Assessment

1. Write the general formula for alkenes and explain the meaning of each term in it.
2. Give the molecular formula and name of the alkene containing four carbon atoms.
3. State the general formula of alkynes and explain what it indicates about the number of hydrogen atoms.
4. Differentiate between alkenes and alkynes based on their types of carbon–carbon bonds.

12.7 Isomerism in Organic Compounds

Compounds with the same molecular formula but different structural formulas are known as isomers, and the phenomenon is referred to as isomerism. For example, ethyl alcohol and dimethyl ether have the same molecular formula (C_2H_6O), but different structural formulae.



Ethyl alcohol Dimethyl ether

Fig. 12.3: Isomerism in organic compounds

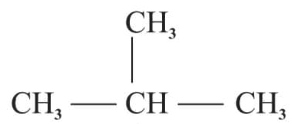
The number of isomers depends on the number of carbon atoms in the molecule. The number of isomers increases as the number of carbon atoms in the molecule increases.

12.7.1 Structural Isomers

Compounds with the same molecular formula but different structural formulas are known as structural isomers, and the phenomenon is termed structural isomerism. Structural isomers differ in the arrangement of atoms within the molecule. For example, n-butane and isobutane are isomers; they share the same molecular formula (C_4H_{10}) but have distinct structures.

$$CH_3 - CH_2 - CH_2 - CH_3$$

n-Butane



Isobutane

Additionally, 1-butene and 2-butene are structural isomers; they share the same molecular formula (C_4H_8) but differ in the positioning of their functional group (the double bond).

$$CH_3 - CH_2 - CH = CH_2$$

1-Butene

$$CH_3 - CH = CH - CH_3$$

2-Butene

Isomers have different physical, chemical, and biological properties.

ACITIVITY

Use molecular model kits or coloured beads to construct models of butane and isobutane to visualise structural isomerism.

12.8 Alkanes

Hydrocarbons with single bonds between their carbon atoms are known as saturated hydrocarbons. They are also referred to as paraffin. The term paraffin is derived from the Latin words "parum", meaning "little", and "affin", meaning "affinity". Hence, the combination of these words means "little affinity" or "the least reactive".

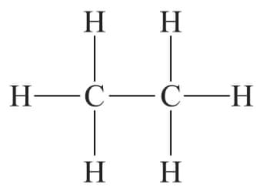


Fig. 12.4: Two dimensional representation of ethane molecule

Alkanes are saturated hydrocarbons in which the four valencies of carbon are fully satisfied; hence, no further atom or group of atoms can bond to carbon atoms without the removal of hydrogen atoms. They do not have a functional group. Examples include methane, ethane, propane, and butane. Their general formula is C_nH_{2n+2}.

Table 12.7: Some saturated hydrocarbons

Saturated Hydrocarbons	Line formula
Methane	CH_4
Ethane	$CH_3 - CH_3$
Propane	$CH_3 - CH_2 - CH_3$
Butane	$CH_3 - CH_2 - CH_2 - CH_3$

CHEMISTRY IN LIFE

Alkanes are used to produce liquefied petroleum gas (LPG) and petrol that power our vehicles.

12.9 Preparation of Alkanes

12.9.1 Cracking of Larger Hydrocarbons

Cracking is a process where larger hydrocarbon molecules are broken down into smaller, more useful hydrocarbons, such as alkanes and alkenes. This process occurs by heating heavy fractions of petroleum either in the presence of a catalyst (catalytic cracking) or without a catalyst (thermal cracking).

Cracking is an essential industrial process used in petroleum refining to produce smaller hydrocarbons such as methane, ethane, propane, and butane, which are valuable as fuels and feedstocks in the chemical industry.

i) Thermal Cracking: In this method, high-molecular-weight hydrocarbons are heated to extremely high temperatures (approximately 700–900°C) in the absence of oxygen. The molecules break down into smaller hydrocarbons.

$$\underset{\text{Decane}}{C_{10}H_{22}} \xrightarrow{\text{Heat}} \underset{\text{Octane}}{C_8H_{18}} + \underset{\text{Ethylene}}{C_2H_4}$$

ii) Catalytic Cracking: This process occurs at lower temperatures (450–550°C) in the presence of a zeolite or aluminosilicate catalyst. It produces a higher yield of branched and cyclic alkanes that are important components of petrol.

$$\underset{\text{Hexadecane}}{C_{16}H_{34}} \xrightarrow[500^\circ C]{\text{Zeolite}} \underset{\text{Octane}}{C_8H_{18}} + \underset{\text{Butene}}{2C_4H_8}$$

Real-Life Application

Catalytic cracking in refineries helps convert heavy oil fractions into petrol and diesel.

12.9.2 Hydrogenation of Alkenes and Alkynes

Hydrogenation refers to the addition of hydrogen. As alkenes and alkynes are deficient in hydrogen, they can form corresponding alkanes when we add the necessary number of hydrogen atoms. This is achieved by reacting the alkene and alkyne with hydrogen in the presence of a catalyst, typically Nickel, Platinum, or Palladium. This process of adding hydrogen is termed hydrogenation. For example:

$$\underset{\text{Ethene}}{CH_2=CH_2} + H_2 \xrightarrow[\text{or Pt/Pd (25°C)}]{\text{Ni (200-250°C)}} \underset{\text{Ethane}}{CH_3-CH_3}$$

The process of preparing alkanes from alkynes is similar, yet it occurs in two steps. The first step involves preparing an alkene, which is then converted into the corresponding alkanes using the same method described earlier. That is:

Step 1: $$\underset{\text{Ethyne}}{CH\equiv CH} + H_2 \xrightarrow[25^\circ C]{Pd} \underset{\text{Ethene}}{CH_2=CH_2}$$

Step 2: $$\underset{\text{Ethene}}{CH_2=CH_2} + H_2 \xrightarrow[25^\circ C]{Pd} \underset{\text{Ethane}}{CH_3-CH_3}$$

12.9.3 Reduction of Alkyl Halides

Another method of preparing alkanes is the reduction of alkyl halides. Alkyl halides are compounds in which a halogen atom is present in the alkyl chain. They are relatively more reactive, making this reaction easy to perform without the need for any catalyst or excessive heat; room temperature is sufficient. Only a suitable reducing agent, such as zinc metal, is required. For example:

$$\underset{\text{Ethyl Chloride}}{H_3C-CH_2-Cl} + 2[H] \xrightarrow{Zn/HCl} \underset{\text{Ethane}}{H_3C-CH_3} + HCl$$

Interesting Information

The solid alkanes are used in waxy coatings added to fruits (apple, plums, oranges, pears, etc). and vegetables (cucumber, turnips, green tomatoes etc.) to:

- Prevent loss of water that helps to maintain firmness and juiciness.
- Improve appearance and increase visual freshness
- Slow down the natural degradation by microbes.

12.10 Chemical Properties of Alkanes

Alkanes are generally regarded as chemically unreactive, except in combustion and chlorination reactions. This is due to the stable, non-polar nature of the carbon-carbon and carbon-hydrogen bonds in alkanes. The electrons in the C-H and C-C sigma bonds are shared equally by the bonding atoms, indicating that the atoms do not have a significant charge.

Interesting Information

Gas lighters utilize liquid alkane (propane or butane) molecules as fuel, providing convenience for lighting candles, stoves and grills.

12.10.1 Combustion

The combustion of alkanes occurs when they burn in the presence of oxygen. Alkanes burn readily in oxygen, producing carbon dioxide, water vapour, and a significant amount of heat energy.

i) Complete Combustion: The general equation for the complete combustion of alkanes is:

$$\text{Alkane}_{(g)} + \underset{\text{(Excess)}}{O_{2(g)}} \xrightarrow{\text{Flameor Ignition}} CO_{2(g)} + H_2O_{(g)} + \text{Energy}$$

Methane gas is found in natural gas and used for cooking and heating. The equation for the combustion of methane is written as:

$$CH_{4(g)} + 2O_{2(g)} \xrightarrow{\text{Flameor Ignition}} CO_{2(g)} + 2H_2O_{(g)} + \Delta H = -890.3\text{kJ/mol}$$

The heat of combustion is defined as the amount of heat released when one mole of a hydrocarbon is completely combusted. During this process, the carbon atoms in the alkane are oxidised to carbon dioxide, and the hydrogen atoms are oxidised to water, liberating a large quantity of energy. This makes alkanes valuable as fuels.

AMAZING FACT The combustion of one mole of methane releases nearly 890 kJ of energy, enough to heat about 30 litres of water from 25°C to boiling point.

ii) Incomplete Combustion: The incomplete combustion of alkanes occurs when there is an insufficient supply of oxygen, resulting in the formation of carbon monoxide, carbon soot, and water vapours.

$$3CH_{4(g)} + 4O_{2(g)} \xrightarrow{\text{Flame or Ignition}} 2CO_{(g)} + C_{(s)} + 6H_2O_{(g)}$$

This process is less efficient and produces harmful by-products.

Real-World Connections

Incomplete combustion of alkanes in car engines produces toxic carbon monoxide, a serious air pollutant.

iii) Applications: Alkanes are widely used as fuels due to their low cost and high availability. They can be easily extracted from crude oil and natural gas. They produce a significant amount of heat per unit mass, and their combustion can be easily regulated. Lighter alkanes, such as methane, propane, and butane, are particularly advantageous for cooking, heating, and power generation due to their high energy content and ease of use.

Do You Know?

Methane is approximately 25 times more efficient than carbon dioxide at trapping heat in the atmosphere.

12.10.2 Substitution

Halogenation is a fundamental reaction in which alkanes undergo substitution by halogen atoms. This process involves the replacement of one or more hydrogen atoms in the alkane molecule with halogen atoms, typically chlorine or bromine. The reaction is usually initiated by the presence of ultraviolet light or elevated temperatures, which provide the necessary energy to break the carbon-hydrogen bonds and allow the halogen atoms to position themselves at specific sites within the alkane. For instance, alkanes react with halogens (chlorine atoms) in the presence of light to produce alkyl halides.

$$CH_{4(g)} + Cl_{2(g)} \xrightarrow{\text{UV Light}} CH_3Cl_{(g)} + HCl_{(g)}$$

$$CH_3CH_{3(g)} + Cl_{2(g)} \xrightarrow{\text{UV Light}} CH_3CH_2Cl_{(g)} + HCl_{(g)}$$

In this reaction, one or more hydrogen atoms in the alkane molecule are substituted by chlorine atoms, resulting in chloroalkanes (also known as alkyl halides).

KEY POINTS

1. Hydrocarbons, made up only of carbon and hydrogen, are the simplest organic compounds and serve as key components of fuels and industrial raw materials.
2. Saturated hydrocarbons only have single bonds, whereas unsaturated hydrocarbons contain double or triple bonds.
3. Isomers share the same molecular formula but differ in their structural arrangements, leading to unique properties.
4. Alkanes (paraffins) follow the general formula C_nH_{2n+2} and include methane, ethane, propane, and butane.
5. Alkanes are prepared by cracking of larger hydrocarbons, hydrogenation of alkenes and alkynes, and reduction of alkyl halides.
6. Alkanes are relatively inert because of strong non-polar C—H bonds and undergo mainly substitution and combustion reactions.
7. Organic compounds from plants include carbohydrates, vitamins, alcohols, and acids, while those from animals include proteins, hormones, fats, and urea.
8. Functional groups are specific atoms or groups of atoms that determine the chemical behaviour of a molecule; compounds with the same functional group show similar properties.
9. A homologous series is a family of compounds with similar structures and the same functional group, where each successive member differs by a methylene group ($—CH_2—$).

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The compound that is **NOT** considered organic is:
(a) methane (b) ethanol
(c) sodium carbonate (d) acetic acid

ii) The general formula of alkanes is:
(a) C_nH_{2n} (b) C_nH_{2n-2}
(c) C_nH_{2n+2} (d) $C_nH_{2n+1}OH$

iii) The unsaturated hydrocarbon among the following is:
(a) methane (b) ethane
(c) ethene (d) propane

iv) The type of hydrocarbon that is C_4H_6 is:
(a) alkane (b) alkene
(c) alkyne (d) aromatic compound

v) The true statement about homologous series is:
- (a) members differ by a –COOH group
- (b) all have different functional groups
- (c) each member differs by a $-CH_2-$ unit
- (d) all have the same molecular mass

vi) The cyclic hydrocarbon among the following is:
- (a) butane
- (b) propane
- (c) benzene
- (d) methane

vii) The catalyst commonly used in the hydrogenation of alkenes and alkynes is:
- (a) iron
- (b) nickel
- (c) copper
- (d) cobalt

viii) The product of the complete combustion of methane in the presence of oxygen is:
- (a) CO_2
- (b) CO
- (c) CH_4
- (d) C_2H_6

ix) The reaction that involves the replacement of hydrogen atoms in alkanes by halogen atoms is:
- (a) hydrogenation
- (b) dehydration
- (c) halogenation
- (d) oxidation

x). The primary product when ethane undergoes complete combustion is:
- (a) carbon monoxide and water
- (b) carbon dioxide and water
- (c) carbon soot and water
- (d) ethylene and hydrogen

B. Restricted Response Questions (RRQs)

i) Name four elements commonly found in organic compounds besides carbon.

ii) Give one example each of a straight-chain, branched, and cyclic organic compound.

iii) What does a structural formula represent?

iv) What is the difference between two consecutive members of a homologous series?

v) Which functional group is common in methanol and ethanol?

vi) Do compounds with the same functional group have similar chemical properties?

vii) Write the balanced chemical equation for the complete combustion of methane.

C. Extended Response Questions (ERQs)

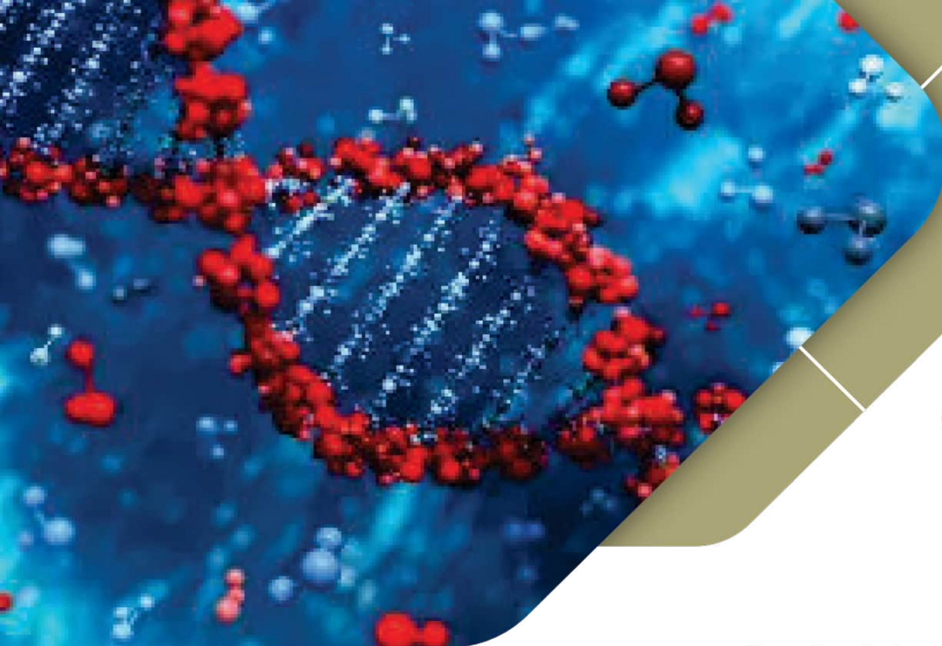
i) Describe the concept of a homologous series. What are the general characteristics of homologous series?

ii) What are structural isomers? Give two pairs of examples.

iii) Compare and contrast alkanes, alkenes, and alkynes in terms of their general formulae, types of bonds, and examples.

iv) Compare and contrast complete and incomplete combustion of alkanes. Include chemical equations.

v) Discuss the substitution reaction of alkanes with halogens. What conditions are required for this reaction?



UNIT 13

BIOCHEMISTRY

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

E-17 Explain the importance and basics of nutrition and healthy eating.

E-18 Recognize the main biomolecules; carbohydrates, proteins, lipids and nucleic acids. Their sources, along with the required daily intake for young adults.

E-19 Identify carbohydrates as a source of energy.

Introduction

Chemistry is the study of matter. A large portion of our world consists of matter we call living organisms. The matter that living bodies consist of is studied in biochemistry. Biochemistry is the branch of chemistry that deals with the composition, structure, and properties of matter found in living organisms.

Living organisms show considerable diversity in their behaviour, structure, and form. However, they are primarily composed of the same types of elements. These types significantly help in understanding life. The most common compounds found in living organisms include carbohydrates, proteins, lipids, nucleic acids, and vitamins.

The principles established by biochemistry help us to improve our quality of life and combat diseases by developing and understanding the systems within living organisms. This includes the development of new and superior crop varieties through genetic engineering techniques, among many other applications.

13.1 Carbohydrates

Carbohydrates are biomolecules composed of carbon, hydrogen, and oxygen atoms, typically in a 1:2:1 ratio, represented by the general formula ($C_n(H_2O)_n$). They are the most abundant class of biomolecules on Earth and serve as a primary energy source for most living organisms. Carbohydrates are essential for various biological functions, including providing energy, storing energy, and forming structural components of cells and tissues.

Interesting Information

Photosynthesis is the vital process by which plants convert over 100 billion tonnes of carbon dioxide and water into carbohydrates annually, sustaining life on Earth and producing the oxygen we breathe. This process is essential for the production of glucose, a primary source of energy for both plants and animals.

13.1.1 Main Classes of Carbohydrates

Carbohydrates are classified into three main groups based on the number of sugar units present: monosaccharides, oligosaccharides, and polysaccharides.

Carbohydrates

Monosaccharides	Oligosaccharides	Polysaccharides
Simple sugars **Examples:** 1. Galactose 2. Fructose 3. Glucose	On hydrolysis yield 2 -10 monosacchairdes. **Examples:** 1. Sucrose 2. Lactose 3. Raffinose	Complex and abundant in nature. **Examples:** 1. Starch 2. Glycogen 3. Cellulose

Do You Know?

The human brain alone uses about 120 grams of glucose daily, making up nearly 60% of the body's total glucose consumption.

13.1.2 Sources of Carbohydrates

They can be found in a wide variety of foods, including both natural and processed options. The primary dietary sources of carbohydrates consist of cereal grains (wheat, rice, maize, barley, oats), starchy vegetables (potatoes, sweet potatoes, corn, peas), fruits (bananas, apples, oranges, grapes), legumes (beans, lentils, peas), and dairy products (milk, yoghurt).

Fig.13.1: Natural sources of carbohydrates

Table 13.1: Some essential carbohydrates and their sources

Carbohydrate	Sources
Glucose	Fruits, vegetables, and honey
Fructose	Fruits, honey, and some vegetables
Sucrose (Table Sugar)	Sugarcane, sugar beet, fruits, and vegetables
Lactose	Milk and other dairy products
Starch	Potato, cereals like wheat, rice, and maize
Glycogen	Animal liver

13.1.3 Recommended Daily Intake of Carbohydrates

The recommended daily intake of carbohydrates for young adults is 225–325 grams, making up 45–65% of a 2,000-calorie diet. The exact amount may vary based on an individual's activity level, metabolism, and overall health goals. Furthermore, the World Health Organisation (WHO) recommends that the daily caloric intake from sucrose should not exceed 5%, as excessive consumption of added sugars can lead to obesity and dental caries

Interesting Information

Importance of Fibre in the Diet

The recommended daily fibre intake is 21–25 grams for women and 30–38 grams for men to support healthy digestion and overall well-being. Fibre is essential for maintaining regular bowel movements and can be found in a variety of foods, such as fruits, vegetables, whole grains, and legumes.

13.1.4 Carbohydrates as a Source of Energy

Carbohydrates are a vital source of energy for the human body, providing about 4kcal/g. When consumed, carbohydrates are broken down into simple sugars, mainly glucose, by digestive enzymes. The small intestine absorbs glucose directly into the bloodstream, which then transports it to various body cells, where it is used to produce energy through respiration.

Excess glucose is converted into glycogen and stored in the liver and muscles. This glycogen acts as a short-term energy reserve; when energy is required, it is converted back into glucose. If the body does not utilise all the glucose for energy, it can be converted into fat and stored in adipose tissue. This process can contribute to obesity if it is not balanced with adequate physical activity and a healthy diet.

Interesting Information

The carbohydrates are the instant source of energy: In hospitals 5% glucose solution marketed as 5% dextrose is used frequently in the form of drips for parenteral nutrition to treat the patients who can not take food or have some problems in eating food.This saves lives of thousands of humans:

13.2 The Proteins

The word "protein" originates from the Greek term "proteios," meaning "first.". In the human body, proteins provide structure and support through skin, hair, cartilage, muscles, tendons, and ligaments. They also function as enzymes, hormones, and antibodies that catalyse, regulate, and protect body chemistry. Additionally, proteins such as haemoglobin, myoglobin, and lipoproteins facilitate the transport of oxygen and other substances within the body. Proteins are polymers of amino acids linked by peptide bonds (—CO—NH—). The human body uses various amino acids to form proteins, with ten being synthesised by our body and the remaining ten obtained from food sources. The amino acids that are essential in our diet are known as essential amino acids.

13.2.1 Sources of Proteins

Proteins are vital to every living cell and are found in all forms of life. Plants can produce their own proteins using simple elements, but animals mainly obtain their proteins by consuming amino acids from their food. However, they can synthesise some amino acids on their own.

We can obtain the required proteins or amino acids from various sources, such as meat, eggs, beans, fruits, lentils, pulses, dried fruits, and seeds.

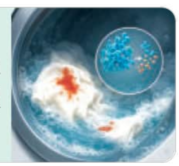
Fig. 13.2: Natural sources of proteins

13.2.2 Recommended Daily Intake

The recommended daily protein intake for young adults is approximately 46–56 grams, making up about 10–15% of total daily calories. The exact amount varies based on factors such as age, gender, body weight, level of physical activity, and overall health.

Real-Life Application

Enzymes obtained from bacteria are commonly used in modern detergents to remove protein-based stains, including blood, egg, and sweat, from clothes.



13.3 The Lipids

Lipids are esters of fatty acids and glycerol. They form a vital class of biomolecules that are found across all living organisms. The most well-known lipids are fats and oils, which are triesters of glycerol (also known as triglycerides) and long-chain carboxylic acids, commonly referred to as fatty acids.

Chemistry in Life

The lipid layer beneath our skin helps insulate the body against temperature changes.

Other important types of lipids include phospholipids, steroids, terpenes, and waxes.

Lipids are generally represented by the structural formula shown in Figure 13.3, where R_1, R_2, and R_3 represent hydrocarbon chains of different lengths derived from fatty acids.

Fats, among other essential components, are an integral part of the human diet, providing a significant source of energy. Humans consume animal fat, margarine, butter, and cream. The oils are unsaturated compounds that are liquid at normal temperatures. They are usually found in plant products.

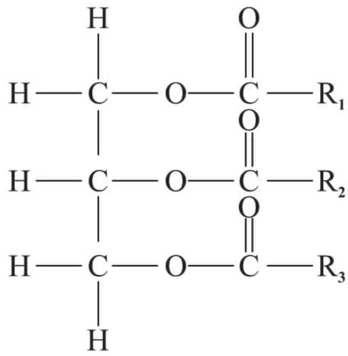


Fig. 13.3: General formula of lipid

Interesting Information

Beeswax is a natural substance produced by bees to construct hive cells for honey and brood (bee eggs). In contrast, paraffin wax, commonly used in candle production, is a petroleum-derived mixture of high molecular weight alkanes. Unlike beeswax, paraffin wax does not contain the ester functional group. This distinction highlights the unique composition and origin of beeswax compared to synthetic waxes like paraffin.

13.3.1 Sources of Lipids

Lipids are obtained from both the plants and animals. In animals, they are mainly found as fats, whereas in plants, they are found as oils. Common animal sources of lipids include butter, lard, tallow, fish oil, and egg yolk. Major plant sources of oils include olive oil, coconut oil, sunflower oil, soybean oil, mustard oil, and peanut oil (also known as groundnut oil). These lipids act as concentrated energy sources and are essential for cell structure, insulation, and hormone production.

Fig.13.4: Natural sources of lipids

13.3.2 Recommended Daily Intake

Young adults should obtain about 20–35% of their daily calories from fats (45–80 g per day).

MISCONCEPTION Vs REALITY

Myth: All fats are harmful to health.

Reality: Not all fats are harmful; healthy fats like omega-3 and omega-6 are vital for proper brain function, heart health, and hormone regulation.

13.4 Nucleic Acids

Nucleic acids are essential biopolymers found in all living organisms. They are responsible for heredity and cellular functions. The two main types of nucleic acids are DNA (Deoxyribonucleic Acid) and RNA (Ribonucleic Acid). These molecules store and transmit genetic information and direct the synthesis of proteins necessary for all cellular activities.

13.4.1 Deoxyribonucleic Acid (DNA)

DNA is the molecule that defines genetic identity and regulates cellular functions. It consists of repeating units called nucleotides, each consisting of a phosphate group, a deoxyribose sugar, and one of four nitrogenous bases: adenine (A), thymine (T), cytosine (C), and guanine (G). The sequence of these bases forms the genetic code. DNA is a double-stranded helix, characterised by specific base pairing (A with T, and C with G). It provides the instructions for protein synthesis, which dictate how cells behave and how tissues perform distinct functions, despite originating from the same cell.

Interesting Information

You may have heard about DNA tests. The coding in DNA is not only different in various organisms, but each human being has their own characteristic sequence inherited from their parents. In a DNA test, this characteristic sequence is matched, and hence, many problems can be resolved.

13.4.2 Ribonucleic Acid (RNA)

RNA is a single-stranded nucleic acid consisting of nucleotides with ribose sugar instead of deoxyribose. Moreover, uracil (U) replaces thymine. RNA plays various roles in the cell, most notably as a messenger between DNA and the protein synthesis machinery. It assists in converting the genetic code into proteins and regulates the chemical reactions within cells.

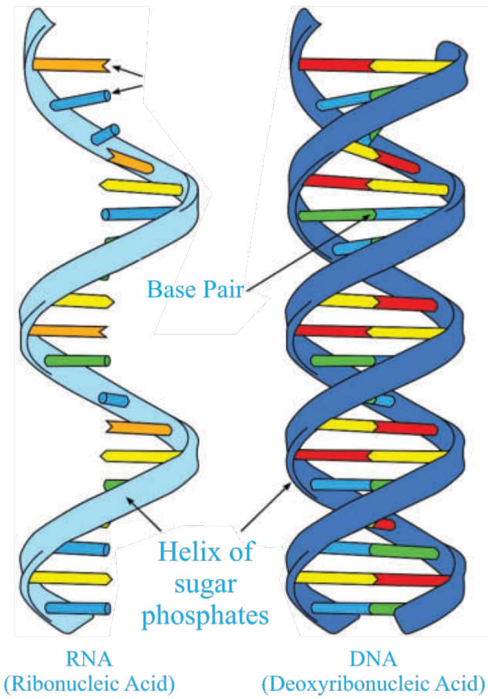


Fig 13.5 Structure of DNA and RNA

13.4.3 Sources of Nucleic Acids in Diet

The body can synthesise nucleotides; however, dietary intake becomes increasingly important during growth, illness, or recovery. Foods rich in nucleic acids include both animal and plant-based choices.

i) Animal Sources: Organ meats, including liver and kidney, as well as meat, fish, and seafood, are rich in DNA and RNA due to their high cellular content.

ii) Plant Sources: Legumes such as beans and lentils, green vegetables like spinach and broccoli, mushrooms, and whole grains provide nucleotides in lower, yet beneficial, amounts.

iii) Other Sources: Yeast and yeast extracts are very high in RNA.

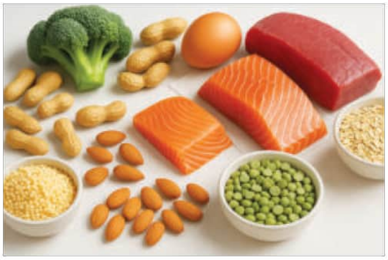


Fig.13.6: Natural sources of nucleic acids

13.4.4 Required Daily Intake

There is no official Recommended Dietary Allowance (RDA) for nucleic acids; however, adults typically consume 300–600 mg of nucleotides daily through a balanced diet. Young adults may benefit from a daily intake of 500–700 mg to support growth, immunity, and tissue repair. The requirement per kilogram of body weight is higher for infants and young children.

ACITIVITY

List common foods in your kitchen and identify their main nutrient as carbohydrate, protein, or lipid.

13.5 Recommended Daily Intake of Macronutrients for Adults

The general recommendations for daily intake of macronutrients for adults are presented in Table 13.2. It is essential to recognise that individual nutritional needs can vary significantly based on various factors, including age, sex, level of physical activity, and health conditions. The values are expressed in grams per day.

Table 13.2: The general recommendations for daily intake of macronutrients for adults.

Nutrient	Recommended Daily Intake	Functions	Sources
Carbohydrates	45-65% of total calories (225-325g for adults)	Provide energy, store energy, build and repair tissues	Whole grains, fruits, vegetables, legumes
Proteins	10-15% of total calories (46-56g for adults)	Build and repair tissues, make enzymes and hormones, transport molecules	Meat, poultry, fish, eggs, dairy products, legumes, nuts, seeds
Lipids	20-35% of total calories (45-80g for adults)	Store energy, insulate organs, protect nerves, produce hormones	Fatty fish, nuts, seeds, vegetable oils, avocados
Nucleic Acids	Small amount needed	Store and transmit genetic information	DNA and RNA found in all living organisms

13.6 Importance and Basics of Healthy Nutrition and Healthy Eating

Good nutrition and healthy eating habits are essential for maintaining overall health and well-being. They play a crucial role in supporting various physical and mental functions, preventing chronic diseases, and promoting an optimal quality of life.

13.6.1 Importance of Nutrition and Healthy Eating

i) Proper nutrition supplies the body with the essential nutrients required to function at its best. It boosts the immune system, facilitates tissue building and repair, sustains strong bones and teeth, and enhances overall well-being.

ii) For children, adolescents, and pregnant women, proper nutrition is vital for healthy growth, development, and the formation of essential organs and tissues.

iii) Balancing calorie intake with physical activity is essential for maintaining a healthy weight. Healthy eating habits can help prevent weight-related issues and promote a healthy body composition.

iv) A well-balanced diet abundant in a variety of nutrients, including vitamins and minerals, can help prevent chronic diseases such as heart disease, diabetes, certain cancers, etc.

v) Nutrition significantly impacts mental well-being. A diet rich in healthy fats, complex carbohydrates, and essential vitamins and minerals can help enhance cognitive function, regulate mood, and promote overall mental health.

13.6.2 Basics of Nutrition and Healthy Eating

i) Eat various foods from all food groups, including fruits and vegetables, whole grains, lean protein sources (such as fish, beans, poultry, and low-fat dairy products), and healthy fats (like avocados, nuts, seeds, and olive oil).

ii) Be mindful of portion sizes to avoid overeating. Use measuring cups and spoons, pay attention to hunger and fullness cues, and avoid distractions while eating.

iii) Drink plenty of water throughout the day to remain hydrated. Aim for eight glasses of water daily, adjusting according to individual needs and activity levels.

iv) Limit the consumption of processed foods, sugary drinks, and items high in saturated and trans fats. These foods offer minimal nutritional value and may contribute to various health issues.

v) Allow yourself to enjoy treats and less nutritious foods in moderation. The key is to focus on a balanced diet overall and avoid deprivation.

vi) Consume a diverse range of foods within each food group to ensure a broad spectrum of nutrients. This will help ensure you obtain all the essential vitamins, minerals, and other significant dietary components.

vii) When planning your diet, consider individual factors like age, sex, activity level, and health status. Consult a healthcare professional or registered dietitian for personalised advice and guidance.

Interesting Information

In addition to essential nutrients, our diet also requires dietary fibre, which is not digestible. Fibre helps move undigested food through the digestive system and should be consumed in amounts ranging from 20 to 30 grams per day. Sufficient fibre intake promotes digestive health, prevents constipation, and reduces the risk of heart disease, diabetes, weight gain, and certain cancers by lowering blood cholesterol and improving overall gut function. A deficiency in dietary fibre often results in constipation.

KEY POINTS

1. The branch of chemistry that studies the matter found in living organisms is known as biochemistry.
2. In biochemistry, we study carbohydrates, proteins, lipids, nucleic acids, and vitamins.
3. Carbohydrates are chemical compounds that provide instant energy to living organisms when consumed.
4. Carbohydrates are, in fact, ketones or aldehydes that also contain numerous hydroxyl groups.
5. Carbohydrates consist of three types: monosaccharides, oligosaccharides, and polysaccharides.
6. The primary source of carbohydrates is the plant kingdom, which produces them through the process of photosynthesis.
7. Proteins are the biologically organic compounds that constitute the bodily substance of animals. They are formed through the combination of amino acids. Essential sources of protein include meat, eggs, milk, lentils, and other foods rich in protein.
8. Lipids are biomolecules made of fatty acids and glycerol, and we consume them as food.
9. Lipids are present in both animals and plants.
10. The animals have lipids known as fats.
11. In plants, lipids exist in the form of oils, which are in a liquid state at normal temperatures.
12. Lipids provide structure to animals and serve as a blanket to protect them from severe cold.
13. Furthermore, various lipids act as hormones in the form of steroids.
14. Nucleic acids consist of two types: DNA and RNA.
15. DNA is double-stranded and provides instructions to run the cell.
16. RNA is single-stranded and performs various functions, including conveying commands from DNA to a specific site and preparing certain proteins.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The elements present in carbohydrates in a 1:2:1 ratio are:

(a) carbon, hydrogen, oxygen (b) carbon, nitrogen, oxygen
(c) hydrogen, oxygen, nitrogen (d) carbon, hydrogen, nitrogen

ii) The building blocks of proteins are:

(a) monosaccharides (b) fatty acids
(c) amino acids (d) nucleotides

iii) The element present in proteins but absent in carbohydrates is:

(a) carbon (b) hydrogen
(c) nitrogen (d) oxygen

iv) The protein responsible for oxygen transport in the blood is:
(a) myoglobin (b) haemoglobin
(c) keratin (d) collagen

v) Essential amino acids are those that:
(a) the body can synthesise (b) are only found in plants
(c) must be obtained through diet (d) are non-proteinogenic

vi) Lipids are formed by the combination of:
(a) amino acids and nitrogen (b) fatty acids and glycerol
(c) monosaccharides and glucose (d) nucleotides and phosphate

vii) A characteristic of lipids is:
(a) soluble in water (b) composed of amino acids
(c) provide long-term energy storage (d) formed by nucleotide chains

viii) The components of a triglyceride molecule are:
(a) three fatty acids and one glycerol
(b) two glycerols and one fatty acid
(c) three amino acids and one glycerol
(d) one phosphate group and two fatty acids

ix) The nitrogenous base found in RNA but not in DNa is:
(a) thymine (b) adenine
(c) uracil (d) cytosine

x) The primary function of dnA in cells is:
(a) energy storage (b) protein synthesis
(c) genetic information storage (d) catalysing reactions

B. Restricted Response Questions (RRQs)

i) What are the basic units (monomers) of carbohydrates?
ii) Name four food sources rich in protein.
iii) Name the two types of nucleic acids.
iv) Differentiate between fat and oil.
v) What is the recommended daily intake of carbohydrates, proteins, and lipids?
vi) What is the function of enzymes in the body?

C. Extended Response Questions (ERQs)

I) Explain the functions of carbohydrates. Give two examples each of monosaccharides, disaccharides, and polysaccharides.
ii) Describe the structure and dietary sources of lipids.
iii) What are nucleic acids? How do DNA and RNA differ in structure and function?

UNIT 14

EMPIRICAL DATA COLLECTION AND ANALYSIS

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

F-01 Explain that units are standardized for better communication and collaboration, (Some examples may include:
In the field of chemistry, the International System of Units (SI) is used to measure physical quantities such as mass, volume, and temperature. This standardized system ensures that chemists worldwide can use the same units to measure and communicate their results, facilitating communication and collaboration in the field. − Without standardized units, it would be difficult for chemists to compare their results with one another, and it would be challenging to develop consistent and accurate scientific models. For example, imagine if one chemist measured the mass of a substance in grams, while another used ounce. The two measurements would be difficult to compare and combine, potentially leading to inaccurate or inconsistent results.)

F-02 Identify SI units for abstract and physical quantities . (Some examples include mass, time and amount of matter.)

F-03 Apply the concept that units can be combined with terms for magnitude, especially kilo, deci, and milli.

F-04 Justify why chemists use 'cm^3' , 'g' and 's' as more practical units when working with small amounts in lab .

F-05 Explain with examples how different tools and techniques can be used to manage accuracy and precision for inherent errors that arise during measurement.

F-06 Use the standard form $A \times 10^n$ where n is a positive or negative integer, and $1 \leq A < 10$

F-07 Convert quantitative values into and out of the scientific notation form.

F-08 Calculate with values in standard form.

F-09 Identify appropriate apparatus for the measurement of time, temperature, mass and volume, including:

a) stopwatches b) thermometers
c) balances d) burettes
e) volumetric pipettes f) measuring cylinders
g) gas syringes

F-10 Suggest advantages and disadvantages of experimental methods and apparatus.

Introduction

Standard scientific notations consisting of a coefficient and an exponent are essential in scientific research and laboratories for clearly expressing large or small numbers. This notation enhances communication and data comparison among scientists, making it easier to represent numerical values accurately.

Common laboratory instruments are essential for conducting scientific experiments and collecting precise measurements and data. Thermometers measure temperature, balances provide accurate mass determinations, and pipettes facilitate precise liquid volume measurements. These instruments play a crucial role in advancing scientific knowledge by promoting accurate data analysis and enabling discoveries.

14.1 The Importance of Using Standard Units in Chemistry

Standardised units play an important role in chemistry, helping scientists communicate and work together easily worldwide. When everyone uses the same measurement system, such as the International System of Units (SI), they can share data, compare results, and repeat experiments without confusion. This is particularly important in chemistry, where even minor variations in measurements can significantly affect the results.

Examples:

I) When chemists measure mass, they use kilograms or grams. If a chemist in Pakistan uses 0.5 kilograms of salt, a chemist in England knows exactly how much salt to use for the same experiment. If some chemists used pounds and others used ounces, it would be difficult to compare results or accurately repeat the experiment. Using the same unit ensures clarity.

ii) Volume is measured in litres, which is an SI unit. If a chemist writes "1.0 litre of hydrochloric acid," chemists everywhere understand exactly how much is meant. If someone were to use pints or gallons, which vary by country, it would be confusing, and experiments would be difficult to compare.

iii) Temperature is measured in kelvin or degrees Celsius. Indicating that a reaction should occur at 298K (which is equivalent to 25°C) is clear to everyone. This enables scientists to calculate energy changes and compare results accurately.

iv) Chemists also measure the amount of a substance using moles. If a reaction requires 2.0 moles of carbon dioxide, everyone knows precisely how much is needed, because a mole is always the same number of particles (about 6.022×10^{23}).

14.2 SI Units for Physical and Abstract Quantities

Physical quantities can be measured directly and possess a physical presence. They are standardised units defined by the International System of Units (SI) to ensure consistency and accuracy in measurements worldwide. Examples include length, time, and temperature. Abstract quantities, in contrast, cannot be measured directly and lack standardised SI units, as they are often subjective and reliant on interpretation rather than direct measurement. Examples include cubit, handspan, and sundial.

In the past, people used non-standardized, abstract units such as cubit (length of the forearm), foot (length of a foot), and handspan (distance between the tip of the thumb and the tip of the little finger) for length measurement, stone (weight of a specific stone) and bushel (volume of grain) for mass measurement, and sundial (position of the sun) and hourglass (flow of sand) for time measurement. These units were significantly inaccurate, varying widely between individuals and locations, lacking consistency, which caused confusion and difficulties in trade and construction. Different regions had unique abstract units, further complicating cross-country communication.

Scientific knowledge and technological advancements have developed over time, increasing the demand for precise and standardised physical measurements. This progress culminated in the establishment of the SI system in 1960, which provides a standardised set of units for measuring various physical quantities such as length, mass, and time. These units ensure measurement consistency across diverse fields and locations, facilitating clear global communication of measurements. This standardisation is crucial for fair trade and commerce and has also played a vital role in the advancement of science, technology, and global collaboration.

Table 14.1: Physical and abstract quantities with their corresponding units

Physical Quantities	SI Units	Abstract Quantities	SI Units
Mass	Kilogram (kg)	Electric Current	Ampere (A)
Time	Second (s)	Luminous Intensity	Candela (cd)
Amount of Substance	Mole (mol)	Electric charge	Coulomb (C)
Length	Meter (m)	Force	Newton (N)
Temperature	Kelvin (K)	Energy	Joule (J)
		Pressure	Pascal (Pa)

14.3 Combining Units with Magnitude Terms (kilo, deci, milli) for Scaling

Prefixes such as kilo (k), deci (d), and milli (m) are used in the International System of Units (SI) to express quantities more flexibly and efficiently. They help in converting between large and small measurement scales, enabling scientists and students to describe quantities conveniently without writing very large or very small numbers.

The prefix kilo (k) denotes a factor of 1000. It is used to describe larger quantities. For example, one kilogram (kg) equals 1000 grams (g), and one kilometre (km) equals 1000 metres (m). This helps in expressing heavy masses or long distances more conveniently. For example, a bag of rice weighing 5 kilograms has a mass of 5,000 grams, and a distance of 12 kilometres equals 12,000 metres.

FUN FACT

The word "kilo" comes from the Greek word "khilioi," meaning "thousand," and "milli" derives from the Latin "millesimus," meaning "thousandth."

The prefix deci (d) indicates a factor of 0.1 or 10^{-1}, meaning one-tenth of a base unit. It helps express intermediate lengths or volumes smaller than one whole unit. For example, one decimetre (dm) equals 0.1 metre, which is also 10 centimetres (cm). If a notebook is 2 decimetres wide, it measures 20 centimetres.

The prefix milli (m) indicates a factor of 0.001 or 10^{-3}, meaning one-thousandth of a base unit. It is commonly used for very small measurements, particularly in science and medicine. For example, one milligram (mg) equals 0.001 gram (g), and one millilitre (mL) equals 0.001 litre (L). A tablet containing 500 milligrams of a substance has a mass of 0.5 grams.

By combining these prefixes with standard SI units, scientists and students can precisely and uniformly express measurements across a broad spectrum of magnitudes. This system ensures clarity, avoids confusion, and simplifies scientific calculations and data interpretation both in classrooms and laboratories.

Self-Assessment

1. What is the relationship between 1 kilometre and 1 metre?
2. Convert 0.003 kilograms into grams and milligrams using appropriate SI units prefixes.

14.4 Practical Advantages of cm^3, g, and s in Chemistry Labs

In chemistry laboratories, the choice of units plays a critical role, particularly when working with small quantities of substances.

14.4.1 Volume Measurement (cm^3)

The cm^3 is a convenient unit for measuring the volume of liquids. It is especially practical for small quantities, offering a manageable and easily visualised unit for volume measurements.

14.4.2 Mass Measurement (g)

The gram is a standard unit of measurement for mass in the metric system. When working with small amounts of substances, balances and scales calibrated in grams provide reliable mass determinations.

14.4.3 Time Measurement (s)

Seconds are a fundamental unit for time measurements and are vital in various chemical experiments. Timing is essential for controlling reaction rates, monitoring reaction progress, and conducting time-dependent procedures. Using seconds ensures accuracy in time-sensitive processes.

14.5 Tools and Techniques to Manage Accuracy and Precision in Measurement

14.5.1 Understanding Accuracy and Precision in Measurements

Accuracy in measurements refers to how closely the measured value aligns with the true or accepted value. It indicates the degree of correctness in a measurement. **Precision,** on the other hand, refers to the closeness of multiple (two or more) measurements to one another, irrespective of whether they are close to the true value. It indicates the consistency or reproducibility of measurements.

Accuracy and Precision Illustration

A dartboard serves as an effective visual tool for understanding the concepts of accuracy and precision. Darts are thrown at the dartboard, with the bullseye representing the true value.

- **No Accuracy, No Precision:** If the darts land away from the bullseye and far apart from each other, the measurements are neither accurate nor precise.
- **High Precision, Low Accuracy:** If the darts land close to one another but far from the bullseye, the measurements are precise yet inaccurate.
- **Low Precision, High Accuracy:** If the darts are dispersed around the bullseye but are generally close to the centre, the measurements are accurate on average but not precise.
- **High Precision, High Accuracy:** If all the darts land near the centre (bullseye) and close to one another, the measurements are both accurate and precise.

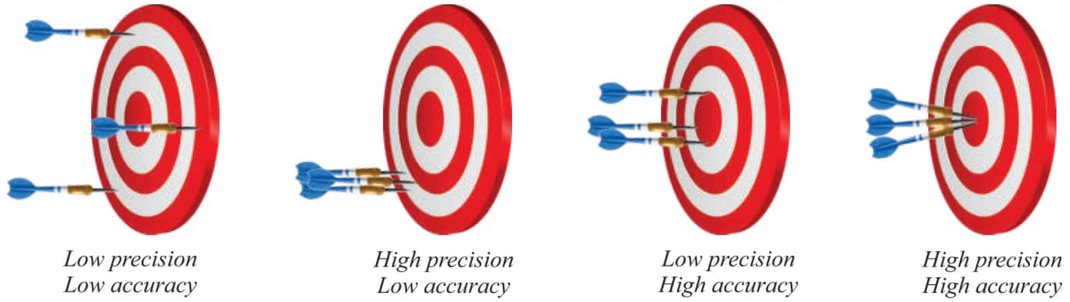


Fig.14.1: Dartboards showing different accuracy and precision situations

14.5.2 Tools for Accurate Measurement

i) Beakers: Beakers hold liquid samples and serve various purposes: containing reactions, stirring, catching filtrates, and mixing. They range from 1mL to several litres, commonly 50mL, 100mL, 500mL, and 1000mL. Volume markings indicate approximate measurements but are not for high-precision tasks.

ii) Graduated Cylinders: They measure the accurate volume of liquids in millilitres, typically from 10mL to 1000mL. They are not used for weighing, mixing, stirring, or heating.

iii) Volumetric Pipettes: These bulb pipettes measure specific liquid volumes, commonly 10mL and 25mL. They ensure accurate and precise volume measurement, making them ideal for preparing solutions and performing titration in labs. A pipette filler draws liquid into the pipette.

iv) Burettes: They are commonly used in titration processes to measure liquid volumes accurately, ranging from 0mL to 50mL. They provide greater accuracy than measuring cylinders but are less precise than pipettes.

v) Gas Syringes: They are used to measure the volume of gases, particularly in chemical experiments. They provide a practical method for collecting and measuring gases produced or consumed during a reaction.

vi) Balances: Balances measure object masses and come in various types such as top-loading and analytical. Top-loading balances provide approximate measurements of large or small samples, while analytical balances accurately measure small masses, often to 0.0001g (0.1mg). These are commonly used in laboratories.

vii) Laboratory Thermometers: These are used to measure the precise temperature of substances.

viii) Stopwatches: Stopwatches precisely measure time. They are frequently used in scientific experiments. They provide an uncomplicated method to measure brief time intervals.

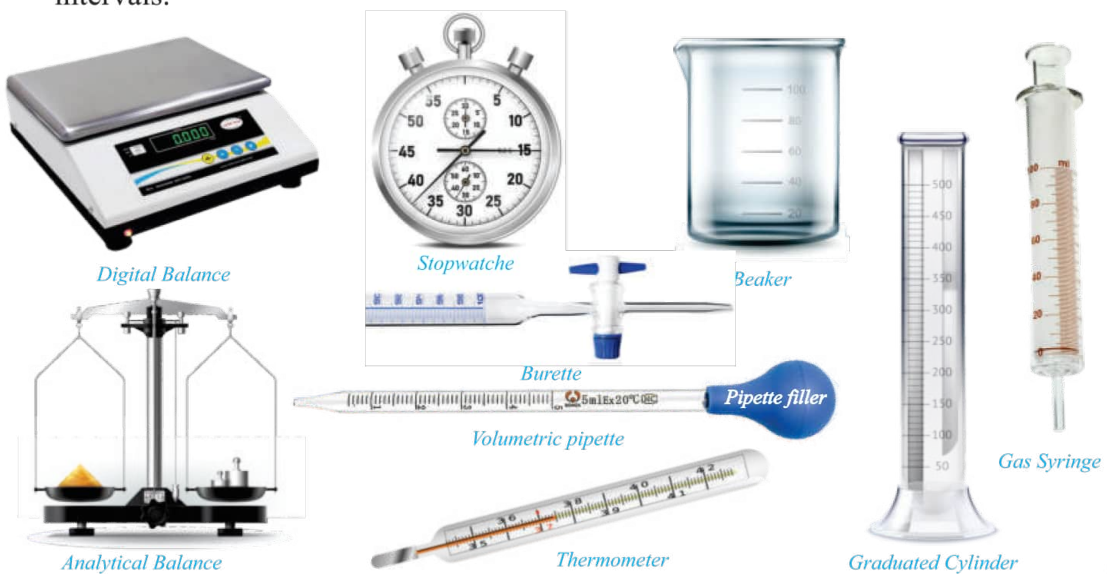


Fig. 14.2: Laboratory equipments

Real-World Connections

In the beverage and perfume industries, volumetric pipettes and burettes are essential for maintaining precise and consistent flavour and fragrance in each batch.

14.5.3 Techniques for Accurate Measurement in A Chemistry Laboratory

i) **Regular Calibration of Equipment:** Regular calibration of instruments such as pipettes, scales, and pH meters is essential for maintaining the integrity of experimental data. For instance, using a graduated cylinder to measure 100 mL of water requires comparing its readings (markings) to a known volume of water, such as a standard beaker.

ii) **Routine Maintenance of Instruments:** Proper maintenance is vital for the accuracy and longevity of laboratory instruments. This may involve routine cleaning, lubricating moving parts, and inspections to prevent deterioration. For example, a well-maintained and dust-free balance for weighing chemicals provides precise weights.

iii) **Choosing the Right Equipment for the Job:** The selection of the right equipment is important for obtaining accurate measurements. When multiple instruments are available, selecting the one with the highest precision that suits the task is essential. For example, using a dropper rather than a graduated cylinder to measure a small quantity of liquid enhances measurement precision.

iv) **Taking Multiple Measurements:** Repeating measurements of the same item can enhance precision and reduce potential errors. For example, performing multiple trials during a titration experiment can lead to a more precise determination of concentration.

v) **Minimising Human Error:** Human error can lead to inaccurate results; therefore, it is important to be attentive and careful when following detailed measurement procedures. Training lab staff to maintain consistent practices, such as pipetting, significantly improves the precision and reliability of their work and reduces human-induced errors.

vi) **Understanding Meniscus Reading:** A meniscus is the curved surface of the liquid in narrow containers, formed due to surface tension. When reading the liquid level, observe the bottom of the meniscus in graduated cylinders for an accurate and precise measurement.

ACITIVITY

Use a beaker, pipette, and burette to measure 25mL of water. Record the readings and discuss which instrument gives the most accurate result.

14.6 Standard Form $A \times 10^n$ for Expressing Numbers

The standard form, expressed as $A \times 10^n$, represents numbers, particularly those that are very large or very small, in a concise manner. In this form, A is a number greater than or equal to 1 and less than 10, while n is an integer that can be either positive or negative.

Coefficient $\longrightarrow A \times 10^n \longleftarrow$ Exponent

Base

14.6.1 Large Number

The value 235,000,000 can be written in standard form as 2.35×10^8. For large numbers, you move the decimal point to the left and count the number of places moved. A becomes 2.35 (between 1 and 10), and n is 8 (positive because the number became smaller).

14.6.2 Small Number

The value 0.0000000000000000000000000000911 can be expressed in standard form as 9.11×10^{-29}. For small numbers, you shift the decimal point to the right and count the number of places moved. This results in 9.11 (which lies between 1 and 10), and n becomes −29 (negative because the number has increased).

Chemistry in Your Body

The human body contains about 3×10^{13} cells, and scientists use scientific notation to express such large numbers easily.

14.7 Conversion of Quantitative Values into and out of Scientific Notation

14.7.1 Conversion of Quantitative Values into Scientific Notation

Avogadro's number (602200000000000000000000) can be easily expressed in exponential form. This is a concise way of representing a large number of zeros. For example, to convert the number 602200000000000000000000 into exponential form, we must shift the decimal point to the right until we arrive at a number between 1 and 10. Count the number of places the decimal point is moved, and that count will become the exponent.

Steps:

1. Move the decimal point to the right 23 places, and we get 6.022.
2. The exponent will be positive since we moved the decimal point to the right.
3. The number 602200000000000000000000 can be written in exponential form as 6.022×10^{23}.

14.7.2 Conversion of Quantitative values out of Scientific Notation

To convert the value 6.625×10^{-34} out of exponential form, we need to shift the decimal point to the left or right according to the value of the exponent. When the exponent is negative, we move the decimal point to the left by the absolute value of the exponent.

Steps:

1. The exponent is −34, indicating that we need to shift the decimal point 34 places to the left.

2. Starting with 6.625, we shift the decimal point 34 places to the left, filling any vacant spaces with zeros.
3. After adjusting the decimal point, we obtain 0.000000000000000000000 0000000000006625 in standard decimal notation.

14.8 Calculations with Values in Standard Form

14.8.1 Multiplication

Exponents are added while multiplying. For example,

$(4.5\times10^{3})(1.2\times10^{5}) \quad = \quad 5.4\times10^{3+5} \quad = \quad 5.4\times10^{8}$

14.8.2 Division

Exponents are subtracted while dividing. For example,

$(3.2\times10^{7})\div(2.5\times10^{3}) \quad = \quad 1.28\times10^{7-3} \quad = \quad 1.28\times10^{4}$

Real-World Connections

Meteorologists measure air pressure in pascals and temperature in kelvin to accurately predict weather patterns and understand atmospheric behaviour with scientific precision.

14.9 Advantages and Disadvantages of Experimental Methods and Apparatus

14.9.1 Advantages of Experimental Methods

Experimental methods are essential in scientific research for studying and understanding various phenomena. These methods provide advantages such as precise control over variables, which allows for the establishment of cause-and-effect relationships. The ability to repeat experiments enhances the reliability of results, thereby increasing confidence in scientific conclusions. Moreover, experiments offer a level of accuracy and understanding that is often unattainable through other research approaches.

14.9.2 Disadvantages of Experimental Methods

Some experimental conditions may not directly apply to real-world situations, as experiments can be resource-intensive, requiring time, funding, and specialised equipment. Not all research questions can be effectively addressed through experimental methods, especially in fields where the complexity of human behaviour or historical context plays a significant role.

14.9.3 Advantages and Disadvantages of Apparatus

1. **Time Measurement (Stopwatches)**

   Advantages: Stopwatches provide precise measurements of time intervals, making them valuable tools for experiments.

   Disadvantages: They have limited precision compared to advanced timing devices and are susceptible to human error.

2. Temperature Measurement (Thermometers)

Advantages: Thermometers are versatile instruments that can measure a wide range of temperatures and are easy to use.

Disadvantages: They are fragile and have limitations in measuring extreme temperatures.

3. Mass Measurement (Balances)

Advantages: Balances provide high-precision measurements for both small and large masses.

Disadvantages: They are sensitive to external factors such as air currents and require careful handling and calibration.

4. Volume Measurement

Burettes: These instruments excel in precise measurement of liquid volumes during titrations. They are highly precise but require careful reading, as they may introduce small errors due to liquid clinging to the walls.

Volumetric Pipettes: They are highly accurate for specific liquid volumes but are fragile and limited to small ones.

Measuring Cylinders: Quick and versatile for rough measurements, but less precise and accurate for small volume readings at the meniscus can be challenging.

Gas Syringes: These are effective in measuring gas volumes but are limited to gases and require careful handling to prevent leakage.

KEY POINTS

1. Standard scientific notation, comprising coefficients and exponents, is essential in research and laboratories for the precise expression of large or small numbers, thereby improving data communication and comparison.
2. Common laboratory instruments, such as thermometers, balances, and pipettes, are essential for ensuring accurate measurements and data collection, thereby advancing scientific knowledge.
3. Standardised units, such as those in the SI system, are essential across various fields, including chemistry.
4. SI units include physical quantities such as mass (kg), time (s), and length (m), as well as abstract quantities like angle (rad) and frequency (Hz), providing a universal language for measurement.
5. Prefixes (kilo, deci, milli) in the SI system facilitate flexible scaling between different measurement scales, ensuring versatility in representing large or small values.

6. In chemistry labs, cm^3, g, and s are preferred for volume, mass, and time, respectively, due to practical advantages like convenience, precision, and compatibility with equipment.
7. Laboratory quality control techniques include calibration, standardisation, quality control charts, and the use of precision instruments and statistical analysis.
8. The standard form ($A \times 10^n$) is used to represent large or small numbers concisely, with examples from astronomical distances and the speed of light.
9. Conversion into and out of scientific notation involves moving the decimal point based on the exponent's sign, ensuring accurate representation of quantities.
10. Laboratory apparatus for measurement includes stopwatches, thermometers, balances, burettes, volumetric pipettes, measuring cylinders, and gas syringes, each serving specific purposes with advantages and disadvantages.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The prefix milli- in the SI system represents:
   (a) one-tenth (b) one-hundredth
   (c) one-thousandth (d) ten thousand

ii) The SI unit of time is:
   (a) minute (b) hour
   (c) second (d) millisecond

iii) The prefix that represents a factor of 0.001 in the SI system is:
   (a) kilo (b) deci
   (c) milli (d) mega

iv) The meaning of the prefix "centi-" in centimetres is:
   (a) 10 (b) 0.1
   (c) 0.01 (d) 100

v) The instrument commonly used to measure small masses is:
   (a) thermometer (b) stopwatch
   (c) analytical balance (d) measuring cylinder

vi) The standard form $A \times 10^n$ represents:
   (a) a large number
   (b) a small number
   (c) both large and small numbers
   (d) neither large nor small numbers

vii) The instrument most suitable for measuring the precise volume of a liquid in a titration is:

(a) beaker (b) graduated cylinder

(c) burette (d) volumetric flask

viii) The SI unit used to measure the amount of substance is:

(a) mole (b) kilogram

(c) ampere (d) litre

ix) The main purpose of using a burette in a laboratory is to:

(a) hold liquids

(b) measure temperature

(c) mix solutions

(d) deliver precise volumes during titration

x) The way exponents are handled in the multiplication of values in standard form is:

(a) added (b) subtracted

(c) multiplied (d) divided

B. Restricted Response Questions (RRQs)

i) Why are standardized units essential in chemistry?

ii) List any three SI base units and their corresponding physical quantities.

iii) Can measurements be precise but not accurate? Explain briefly.

iv) Express 0.00075 in scientific notation.

v) Convert 4.5×10^{6} to standard decimal form.

C. Extended Response Questions (ERQs)

i) Compare physical and abstract quantities, providing two examples of each.

ii) Describe how metric prefixes such as kilo-, deci-, and milli- help express scientific quantities. Provide one example of each.

iii) Explain the difference between accuracy and precision in a chemical measurement. Illustrate your answer with an example.

iv) Explain why scientific notation is important in chemistry. Convert the following number to scientific notation:

(a) 0.0004560mol (b) 4,300,000μL

(c) 0.00089g (d) $560cm^3$

UNIT
15

SEPARATION TECHNIQUES

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

F-11 Define important terms associated with creating chemical solutions. (Some examples include:
- a) solvent as a substance that dissolves a solute
- b) solute as a substance that is dissolved in a solvent
- c) solution as a mixture of one or more solutes dissolved in a solvent
- d) saturated solution as a solution containing the maximum concentration of a solute dissolved in the solvent at a specified temperature
- e) residue as a substance that remains after evaporation, distillation, filtration or any similar process
- f) filtrate as a liquid or solution that has passed through a filter.)

F-12 Explain methods of separation and purification. (Some examples include:
- a) using a suitable solvent (solvent extraction)
- b) filtration
- c) crystallization
- d) simple distillation
- e) fractional distillation
- f) chromatography.)

F-13 Suggest suitable separation and purification techniques, given information about the substances involved, and their usage in daily life.

F-14 Identify substances and assess their purity using melting point and boiling point information.

Introduction

In chemistry, understanding how to prepare solutions and separate their components is important. This foundational knowledge is essential for conducting experiments, analysing substances, and facilitating various chemical reactions and processes in both laboratories and industrial settings. Techniques for separation and purification are vital in pharmaceuticals, environmental science, and food technology. Learning these methods will enable you to isolate specific compounds from mixtures. Moreover, the melting and boiling points of substances serve as key indicators of purity, which play a significant role in quality control across the chemical industry.

15.1 Fundamental Terms in Solution

Solvent: A solvent is a substance, usually a liquid, that can dissolve other substances known as solutes. In a solution, the solvent is the main component and provides the medium in which the solutes are dissolved. For example, water is a common solvent used to dissolve substances such as sugar or salt.

Solute: A solute is a substance that dissolves in a solvent to form a solution. Solutes can exist in various states (solid, liquid, or gas) and are in smaller quantities than the solvent. For example, sugar is a solute when it is dissolved in water.

Solution: A solution is a uniform mixture formed by dissolving one or more solutes in a solvent. In a solution, the solute particles are evenly distributed at the molecular level, resulting in a consistent composition throughout the mixture. For example, saltwater is a solution in which salt (the solute) is dissolved in water (the solvent).

Saturated Solution: A saturated solution contains the maximum quantity of solute that can dissolve in the solvent under specific conditions. Any additional solute introduced will not dissolve and will remain as undissolved particles. For example, when no more salt can dissolve in a glass of water at room temperature, the resulting solution is saturated with salt.

MISCONCEPTION Vs REALITY

Myth: Adding more solute always increases concentration.

Reality: A solution can only contain a fixed amount of solute at a specific temperature. Once it reaches saturation, any extra solute will no longer dissolve and will settle at the bottom as undissolved particles.

Residue: Residue refers to the material that remains after a separation or purification process, such as evaporation, distillation, or filtration. It is the portion that did not undergo the desired transformation or separation and may include undissolved solids. For example, after boiling saltwater until only salt crystals are left, the salt is considered the residue.

Filtrate: Filtrate is the liquid or solution that passes through a filter during the filtration process, leaving behind solid particles or impurities. For example, after filtering a mixture of sand and water, the clear water collected is the filtrate.

Do You Know?

A single teaspoon of seawater contains billions of dissolved ions such as sodium, chloride, and magnesium, which makes it naturally conductive.

15.2 Methods of Separation and Purification

The techniques of separating and purifying substances are vital in chemistry. These techniques allow researchers and industries to produce pure substances for various uses. Each technique uses different physical or chemical properties to reach the aims of separating and purifying the substances.

15.2.1 Using a Suitable Solvent for Selective Dissolution

This method involves selecting a solvent in which the desired substance has high solubility while impurities show low solubility in that solvent . 1st the impure solute is dissolved in a solvent in which solubility of solute is low. Then the mixture is transferred from that solvent into another (selected) solvent in which it is more soluble. The two solvents are immiscible, or at least have very limited miscibility. The phase richer in solute is separated, leaving behind phase having less solute and richer in impurity using a separating funnel.

Fig. 15.1: Separating funnel used for the separation of immiscible liquids

15.2.2 Filtration: Separating by Particle Size

Filtration is a physical method used to separate a solid that does not dissolve in a liquid. In this process, a mixture of the solid and liquid is poured through a filter, typically made of porous materials such as filter paper, cloth, or mesh. The filter allows the liquid component to pass through its small holes or pores, while the undissolved solid particles are too large to pass and become trapped on the surface of the filter. This method is particularly effective for removing insoluble solids like sand from water.

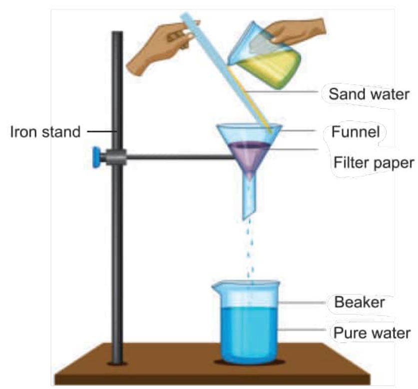


Fig. 15.2: Filtration process for separating sand from water

CONCEPT CHECKPOINT

Why does sand remain on the filter paper during filtration, while salt passes through with the water?

15.2.3 Crystallisation: Purification through Crystal Formation

Crystallisation is a separation and purification technique used to obtain pure solid crystals from a solution. The process begins by preparing a saturated solution at high temperature, meaning that no more solute can dissolve at that temperature. Then it is filtered in hot state to remove undissolved impurities. Then the solution is allowed to cool slowly. As the temperature decreases, the solubility of the solute also decreases, causing the dissolved substance to form solid crystals. These crystals can then be separated from the remaining liquid, which typically contains dissolved impurities. Crystallisation is especially useful for producing highly pure solids from mixtures, as the impurities are removed from the solution while the pure crystals are collected.

Interesting Information

Ancient Egyptians first used crystallisation to purify natural salts from desert lakes.

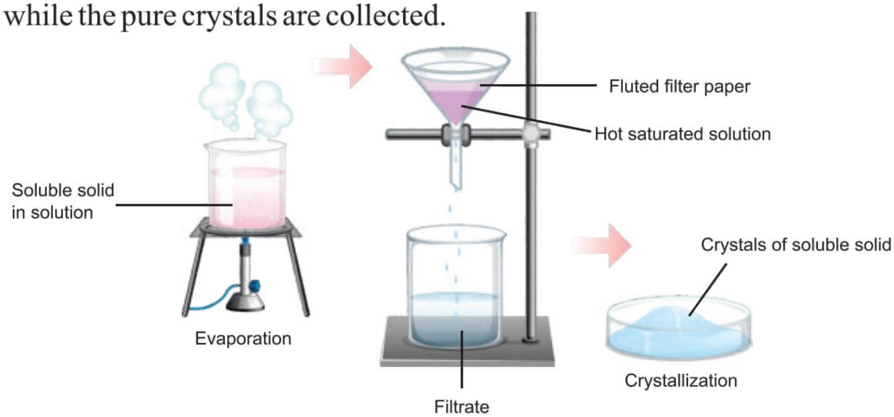


Fig. 15.3: Crystallisation process shows the separation and formation of pure crystals from a hot saturated solution

15.2.4 Simple Distillation: Boiling Point-Based Separation

Simple distillation is a separation technique used to purify a liquid by using the difference in boiling points between substances in a mixture. The mixture is heated in a distillation apparatus, and the component with the lower boiling point vaporises first. This vapour then passes through a condenser, where it cools and turns back into a liquid called the distillate, which is collected separately. Meanwhile, the substance with the higher boiling point remains in the flask.

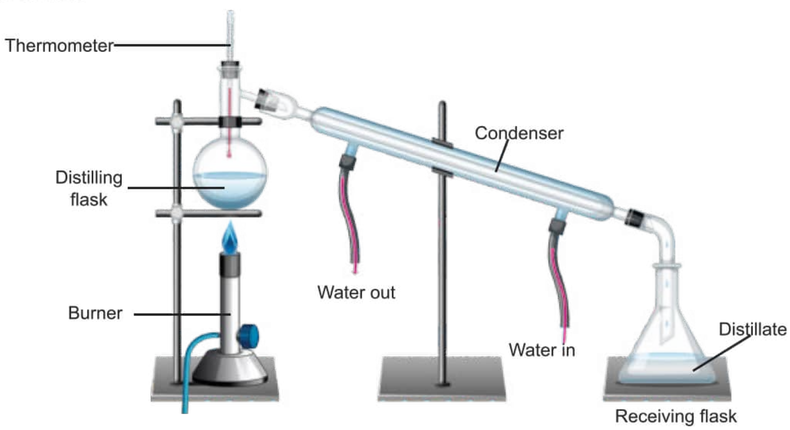


Fig. 15.4: Simple distillation process used to separate a pure liquid from a solution based on differences in boiling points

This method is most effective when the boiling points of the components differ by at least 70°C. Simple distillation is commonly used to obtain pure water from saltwater in desert regions and in the production of alcohol to separate it from mixtures. It is particularly advantageous for removing impurities with higher boiling points or for isolating one liquid from a mixture of liquids.

Real-Life Application

Distillation depends on differences in boiling points to purify water, extract perfumes, and produce fuels.

15.2.5 Fractional Distillation

Fractional distillation is a method used to isolate liquids in a mixture with closely related boiling points. Similar to simple distillation, the mixture is heated, causing the component with the lower boiling point to vaporise first. In contrast to simple distillation, fractional distillation joins a fractionating column situated between the flask and the condenser.

This column enables repeated condensation and vaporisation of the rising vapours, helping to separate the liquids more precisely. Each component rises in the column according to its boiling point: the lower the boiling point, the higher it rises. The vapours are then cooled in the condenser to produce a distillate, which is collected separately.

Fractional distillation is particularly effective when the difference in boiling points is less than 70°C. It is widely used in oil refineries to separate crude oil into valuable products such as petrol, diesel, and kerosene.

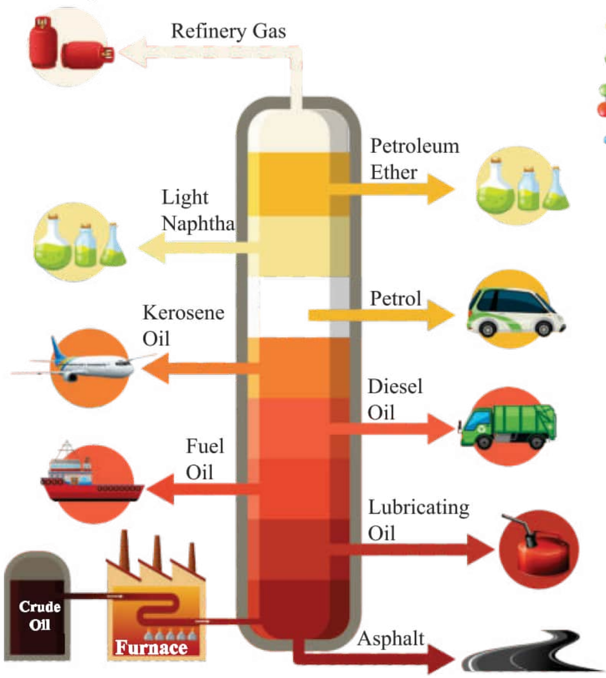


Fig. 15.5: Fractional distillation of petroleum showing the separation of crude oil into useful fractions based on their boiling points

Real-World Connections

Oil refineries separate crude oil into petrol, diesel, and kerosene using large-scale fractional distillation towers.

15.3 Common Separation and Purification Techniques

15.3.1 Filtration: Separation of Salt and Sand Mixture

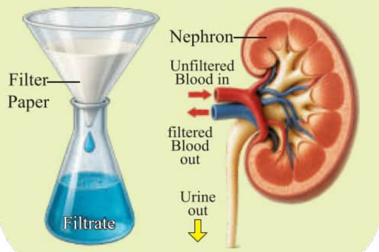
Filtration is a method for separating a mixture of salt and sand. The principle behind filtration is that the solute (in this case, salt) dissolves in a solvent (water), while the insoluble substance (sand) does not.

Procedure

i) Mix the salt and sand with water to dissolve the salt.
ii) The mixture is poured onto a filter paper or filter funnel.
iii) The water containing dissolved salt passes through the filter, leaving the sand as a residue on the filter paper.
iv) The collected filtrate can be further processed through evaporation to obtain pure salt.

Chemistry in Your Body

Your kidneys act like natural filters, removing waste and maintaining ionic balance, similar to filtration in a laboratory.



15.3.2 Decantation: Separating Oil and Water

Decantation is a method used to separate oil and water based on their differing densities and the ability of oil to mix with water. Oil is less dense and does not mix with water, causing it to float on top of the water.

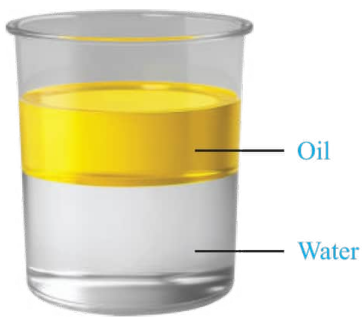


Fig. 15.6: Decantation process showing the separation of oil from water based on their different densities

Procedure

i) Allow the oil-water mixture to stand undisturbed.
ii) The less dense oil forms a distinct layer on top of the water.
iii) Carefully pour off the upper layer containing the oil, leaving the water behind.
iv) This method is simple but may not be suitable for emulsified mixtures where oil and water form a stable suspension

15.3.3 Crystallisation: Purification of Sugar

Crystallisation is a process used to purify sugar. It takes advantage of the fact that substances have different solubilities at different temperatures.

Procedure

i) Dissolve impure sugar in hot water to create a concentrated solution.
ii) Allow the solution to cool slowly. As it cools, sugar molecules come together and form crystals.
iii) The crystals are separated from the remaining liquid by filtration or decantation.
iv) This method effectively obtains high-purity sugar by leaving impurities in the liquid phase.

15.4 Assessing Substance Purity through Melting and Boiling Points

Analysing the melting and boiling points of substances is an important method in chemistry for identifying and assessing their purity. Pure substances exhibit distinct and consistent melting and boiling points, while impurities often lead to deviations or decreased precision in these points.

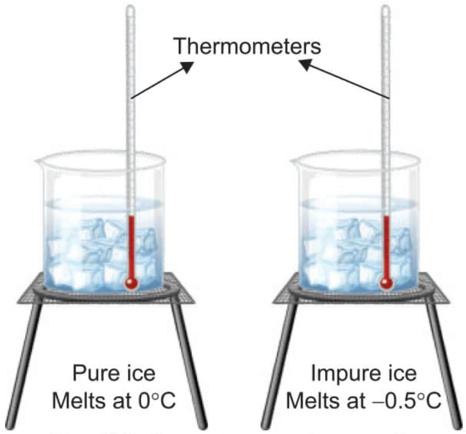


Fig. 15.7: Determination of purity of ice using melting point

15.4.1 Melting Point Analysis

Melting point analysis is a method used to determine the purity of a substance based on its melting characteristics.

i) **Pure Substance Characteristics:** Pure substances have a distinct and sharp melting point, meaning they transition from solid to liquid at a precise temperature. This transition occurs rapidly at the exact melting point.

ii) **Impurity Effects on Melting Point:** Impurities in a substance can lower and broaden the melting point. This is because impurities disrupt the regular crystal lattice structure, making the melting point less precise and causing it to occur over a broader range of temperatures. By comparing the observed melting point to the expected range for a specific substance, qualitative insights into its purity can be gained. A significant deviation or a broader range in the melting point suggests the presence of impurities in the substance.

15.4.2 Boiling Point Analysis

i) **Pure Substance Characteristics:** Pure substances have a specific boiling point, which is the temperature at which they change from liquid to gas. This change occurs quickly and at a particular temperature.

ii) **Impurity Effects on Boiling Point:** Impurities can increase the boiling point and widen the range of temperatures at which the substance boils. Thus, reaching the boiling point requires more energy.
Compare the observed boiling point with the known boiling points of the pure substance as reported in the scientific literature. A significant difference between the observed boiling point and the expected value, or a wider range, suggests the presence of impurities.

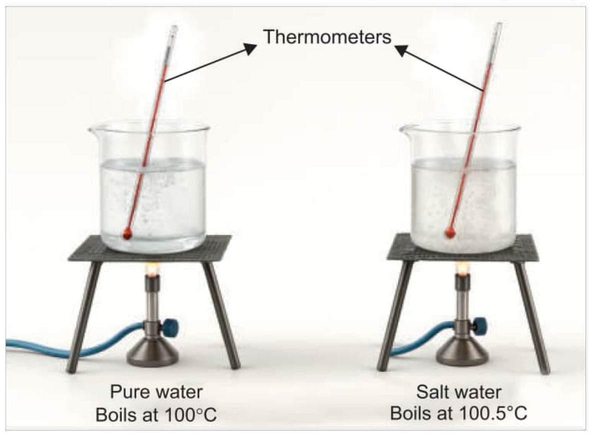


Fig. 15.8: Comparison of boiling points to determine the purity of water

KEY POINTS

1. The solvent is the main component of a solution, usually a liquid, that can dissolve other substances (solutes).
2. A solute is a substance that dissolves in a solvent to form a solution, and is present in a smaller amount than the solvent.
3. A solution is a uniform mixture formed by dissolving one or more solutes in a solvent, leading to an even distribution at the molecular level.
4. A saturated solution holds the maximum amount of solute that can dissolve under specific conditions, with any excess forming undissolved particles.
5. Residue refers to the material left behind after a separation or purification process, such as evaporation or filtration, which may include undissolved solids.
6. Filtrate is the liquid or solution that passes through a filter during filtration, leaving behind solid particles or impurities.
7. Dilution is the process of decreasing the concentration of a solute in a solution by adding more solvent.
8. Concentration is the amount of solute in a specific amount of solvent or solution, usually expressed as molarity.
9. A precipitate is a solid formed when a solute comes out of a solution because of a chemical reaction or a change in conditions.
10. Solubility is the maximum amount of solute that can dissolve in a given amount of solvent at a specific temperature and pressure.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The main component in a solution that provides the medium for dissolving solutes is:

(a) solute	(b) filtrate
(c) solvent	(d) residue

ii) The term for a solution that contains the maximum amount of solute that can dissolve under specific conditions is:

(a) diluted solution	(b) unsaturated solution
(c) saturated solution	(d) concentrated solution

iii) The process of reducing the concentration of a solute in a solution by adding more solvent is:

(a) filtration	(b) dilution
(c) crystallisation	(d) precipitation

iv) The term that describes a solid that forms when a solute comes out of a solution is:
(a) filtrate (b) residue
(c) precipitate (d) solute

v) The method of separation that uses a fractionating column for precise separation of close boiling points is:
(a) simple distillation (b) filtration
(c) crystallisation (d) fractional distillation

vi) The term "miscible" in chemistry refers to:
(a) ability of substances to mix
(b) inability of substances to mix
(c) formation of crystals
(d) presence of solute in a solution

vii) The purpose of a filter in the filtration process is to:
(a) retain solid particles (b) dissolve solutes
(c) increase solubility (d) separate immiscible substances

viii) The method used for separating oil and water based on their difference in density and ability to mix is:
(a) filtration (b) crystallisation
(c) decantation (d) dilution

ix) The purpose of a filter in the filtration process is to:
(a) retain solid particles (b) dissolve solutes
(c) increase solubility (d) separate immiscible substances

x) The term “miscible” in chemistry refers to the:
(a) ability of substances to mix (b) inability of substances to mix
(c) formation of crystals (d) presence of solute in a solution

B. Restricted Response Questions (RRQs)

i) Differentiate between residue and filtrate.
ii) Why does sand remain on the filter paper during filtration?
iii) Why is simple distillation unsuitable for separating liquids with close boiling points?
iv) How do impurities affect the melting point of a substance?
v) What indicates the presence of impurities in boiling point analysis.

C. Extended Response Questions (ERQs)

i) Compare and contrast simple distillation and fractional distillation.
ii) Explain how filtration and decantation differ as separation techniques. Under what conditions is each method most appropriate? Give an example for each.
iii) How can melting point and boiling point analysis be used to assess the purity of a substance?

UNIT 16

QUALITATIVE ANALYSIS

Student Learning Outcomes (SLOs)

After studying this unit, students will be able to:

F-15: Describe tests to identify important gasses.
(Some examples include:
a) ammonia, NH_3, using damp red litmus paper
b) carbon dioxide, CO_2, using limewater
c) chlorine, Cl_2, using damp litmus paper
d) hydrogen, H_2, using a lighted splint
e) oxygen, O_2, using a glowing splint
f) sulphur dioxide, SO_2, using acidified aqueous potassium manganate (VII).)

F-16: Explain the use of a flame test to identify important cations:
(Some examples include:
a) lithium, Li^+
b) sodium, Na^+
c) potassium, K^+
d) calcium, Ca^{2+}
e) copper (II), Cu^{2+}
f) barium, Ba^{2+}.)

F-17 Describe how paper chromatography is used to separate mixtures of soluble substances, using a suitable solvent

F-18 Describe the use of locating agents when separating mixtures containing colourless substances.
(For context, knowledge of specific locating agents is not required

F-19 Interpret simple chromatograms (For context, students should identify:
a) unknown substances by comparison with known substances
b) pure and impure substances

F-20 State and use the equation for R_f.

Introduction

Qualitative analysis involves identifying substances based on their physical and chemical properties, without measuring their quantities. In this unit, you will learn simple tests to identify gases such as ammonia, carbon dioxide, chlorine, hydrogen, oxygen, and sulphur dioxide through distinctive changes in colour, smell, or sound. The flame test identifies metal ions by the characteristic colour they emit when heated in a flame. You will also explore paper chromatography, a method that separates components of a mixture based on their solubility as a solvent moves through filter paper. These techniques are essential tools in qualitative analysis, helping the detection and study of substances in chemistry.

16.1 Tests to Identify Important Gases

16.1.1 Testing for Ammonia Gas (NH_3)

Ammonia gas is produced by heating a mixture of solid ammonium chloride and calcium hydroxide.

$$2NH_4Cl_{(s)} + Ca(OH)_{2(s)} \longrightarrow CaCl_{2(s)} + 2H_2O_{(l)} + 2NH_{3(g)}$$

It is a colourless gas with a sharp, pungent odour and shows basic properties. To test for ammonia, hold a damp red litmus paper near the mouth of the test tube containing the gas, making sure it does not touch the sides to prevent contamination. The litmus paper turns blue, confirming the presence of ammonia gas.

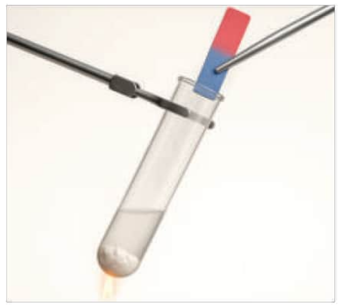


Fig. 16.1: Heating ammonium chloride with calcium hydroxide releases ammonia, which turns damp red litmus paper blue, confirming its basic nature

16.1.2 Testing for Carbon Dioxide Gas (CO_2)

Carbon dioxide gas is produced by reacting calcium carbonate with dilute hydrochloric acid.

$$CaCO_{3(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

To test for carbon dioxide, pass the gas through limewater (an aqueous solution of calcium hydroxide):

$$Ca(OH)_{2(aq)} + CO_{2(aq)} \longrightarrow CaCl_{3(s)} + H_2O_{(l)}$$

The formation of a white precipitate of calcium carbonate turns the limewater milky, confirming the presence of carbon dioxide gas.

Fig.16.2: Carbon dioxide turns clear limewater milky because of the formation of calcium carbonate

16.1.3 Testing for Chlorine Gas (Cl_2)

Chlorine gas is prepared by reacting concentrated hydrochloric acid with manganese dioxide:

$$MnO_{2(s)} + 4HCl \longrightarrow MnCl_{2(aq)} + 2H_2O_{(l)} + Cl_{2(g)}$$

Chlorine functions as a powerful bleaching agent that oxidises coloured substances. To test for chlorine, dampen a small piece of blue litmus paper with a few drops of water and hold it in the gas using tweezers or gloves. Initially, the chlorine gas causes the damp blue litmus paper to turn red due to its acidic nature; subsequently, it bleaches the paper white by removing its colour.

Fig.16.3: Chlorine gas causes damp blue litmus paper to turn red and then bleach it white.

16.1.4 Testing for Hydrogen Gas (H_2)

Hydrogen is a flammable gas that burns in air to produce water.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

To test for hydrogen, hold a burning wooden splint near the mouth of the test tube containing the gas. If the gas is hydrogen, it will produce a "pop" sound.

Fig.16.4: Hydrogen gas produces a faint 'pop' sound when ignited with a burning wooden splint

16.1.5 Testing for Oxygen Gas (O_2)

Oxygen gas supports combustion and acts as an oxidising agent. It can be obtained by the decomposition of hydrogen peroxide in the presence of a catalyst such as manganese dioxide:

$$2H_2O_{2(aq)} \xrightarrow{MnO_2} 2H_2O_{(l)} + O_{2(g)}$$

To test for oxygen, light a wooden splint and gently blow it out until it glows. Hold the glowing splint at the mouth of the test tube containing the gas. If the gas is oxygen, the splint will reignite and burn brightly.

Fig.16.5: A glowing splint reignites when it contacts oxygen gas, confirming its presence.

16.1.6 Identifying Test for Sulphur Dioxide Gas (SO_2)

Sulphur dioxide gas can be identified by the decolourisation of acidified potassium permanganate ($KMnO_4$) solution. Sulphur dioxide reacts with potassium permanganate, and sulphuric acid to produce potassium sulphate, manganese sulphate, and water, resulting in the loss of the solution's purple colour.

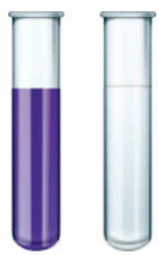


Fig. 16.6: The purple acidified potassium permanganate solution decolourises, confirming the presence of sulphur dioxide

$$5SO_2 + 2KMnO_4 + 2H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 2H_2O$$

16.2 Flame Test

The flame test identifies specific metal cations in compounds by observing the colours they emit when introduced to a flame. Here is how it works and the colours produced by various metal cations:

Procedure: A clean loop of nichrome wire is immersed in the sample compound or salt that contains the metal cation. The wire loop with the sample is then held in the flame of a Bunsen burner or a similar heat source.

Observations: As the metal cation is heated in the flame, it absorbs energy, causing its electrons to become excited. When the excited electrons return to their ground state, they emit light at specific wavelengths, resulting in characteristic colours of the flame.

Table. 16.1: Characteristic flame colours of common metal ions observed during flame tests

Metal Ions	Colour of flame
Lithium (Li^{+})	Crimson red
Sodium (Na^{+})	Intense yellow
Potassium (K^{+})	Lilac (light purple)
Calcium (Ca^{2+})	Orange-red (brick red)
Barium (Ba^{2+})	Pale green (apple green)
Copper (Cu^{2+})	Blue-green (with blue streaks)

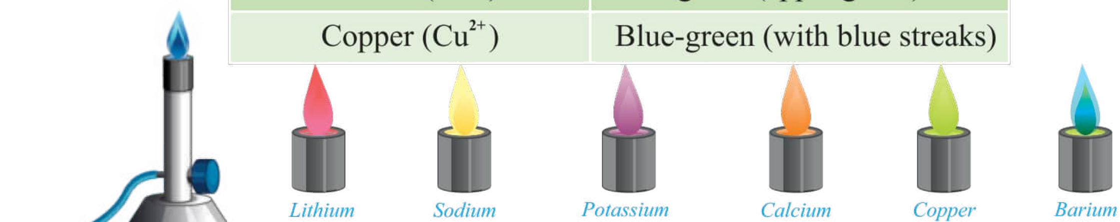


Fig.16.7: Distinctive flame colours produced by various metal ions when heated in a Bunsen flame

The observed flame colour is compared to known standards or reference colours for each metal cation to identify the cation present in the sample.

Observations should be made under controlled conditions to identify the characteristic flame colours accurately.

16.3 Paper Chromatography

Paper chromatography is a simple and effective laboratory technique used to separate and identify the components of a mixture. A mobile phase, usually a solvent, moves through a stationary phase, such as filter paper. Since different substances have different solubilities and affinities for the paper, they travel at different rates, resulting in their separation.

Principle

A piece of filter paper is used as the stationary phase, and a liquid solvent is used as the mobile phase. As the solvent moves up the paper, it carries the different components of the mixture with it. The degree to which each component is carried depends on its solubility in the solvent and its interaction with the paper.

Procedure

The procedure involves preparing the stationary phase by cutting a piece of filter paper and drawing a baseline near one end. The sample to be separated is then spotted onto the baseline using a capillary tube or micropipette. The mobile phase, a solvent mixture such as water and ethanol (or other suitable solvents), is prepared and poured into a container, ensuring that the solvent level remains below the spot on the paper. The paper is then placed into the container, with the lower end immersed in the solvent mixture but not touching the sides of the container. As the solvent moves up the paper, it carries the components of the mixture, separating them. Once the solvent reaches near the top, the paper is removed and allowed to dry. The chromatogram is then analysed by examining it for separated spots and measuring the distance travelled by each spot and the solvent front from the baseline.

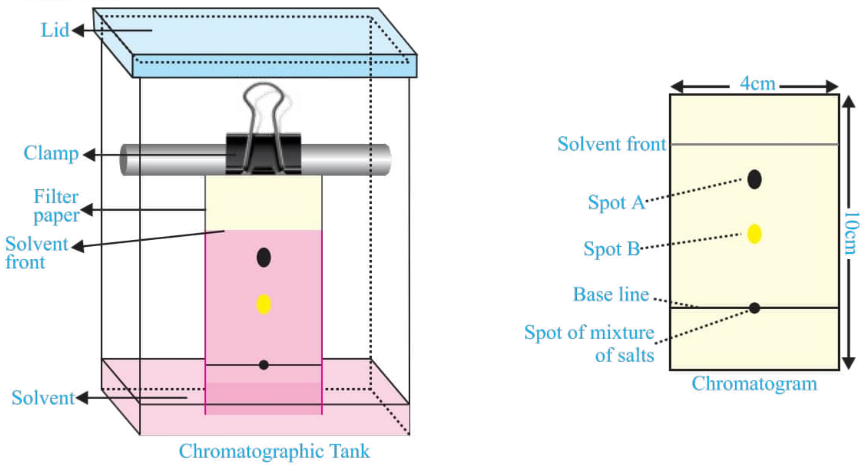


Fig. 16.8: Process of paper chromatography

Keep in Mind:

Solvent Front

The solvent front indicates the highest point the solvent reaches as it rises up the chromatography paper during the experiment. Typically, it is marked after the experiment concludes and signifies how far the solvent has travelled.

Baseline (Origin)

The baseline is the starting line drawn near the bottom of the chromatography paper where the sample mixture is spotted before the experiment begins. It indicates the point from which the substances start to move.

Chromatogram

A chromatogram is the pattern that develops on chromatography paper following the experiment. It shows the separated spots of various components in the mixture and indicates the distance each part has travelled.

16.3.1 Use of Locating Agents

Locating agents are used to make colourless substances visible when separating mixtures. They react with colourless substances to produce coloured or fluorescent products, making them easily visible. For example, iodine vapour can react with fats and oils, causing them to turn brown. Ninhydrin is often used to detect colourless amino acids in paper chromatography.

16.4 Chromatogram Analysis and Identification of Substances

When a chromatogram is produced on filter paper, it shows spots that represent individual components. These spots help us determine which substances are present and whether the sample is pure or a mixture. To analyse a chromatogram, we measure the distance each spot travels from the baseline and calculate a value called R_f (retention or retardation factor), which aids in identifying the substances.

16.4.1 Identification of Unknown Substances

To identify unknown substances, we compare the chromatogram of the unknown sample with those of known substances. If a spot from the unknown appears at the same height and has the same colour as a spot from a known substance, they are likely the same. If the spot is situated at a different position or appears different, it is probably a different substance.

16.4.2 Determination of Purity in Substances

Chromatograms can also indicate whether a substance is pure or impure. A pure substance produces only one spot on the chromatogram, indicating that it contains just one component. On the other hand, an impure substance results in two or more spots, indicating that it contains a mixture of substances.

16.4.3 Calculation and Significance of the R_f Value

The R_f value (retention or retardation factor) is a numerical value used in chromatography to help in identifying substances within a mixture. To determine the R_f value, use this formula:

$$R_f = \frac{\text{Distance traveled by component from baseline}}{\text{Distance traveled by solvent front from baseline}}$$

The R_f value is the ratio of the distance travelled by the component to the distance travelled by the solvent front.

Use of the R_f Equation

Each substance has its own R_f value when the same solvent and paper are used. By calculating the R_f value of an unknown substance and comparing it to known values, you can suggest what the unknown substance might be. If the R_f value of an unknown matches that of a known substance, they could be the same, but it is not always certain because minor differences can occur. Still, R_f values are a useful way to help identify substances in chromatography.

KEY POINTS

1. Ammonia has a pungent smell and turns damp red litmus paper blue. To test for ammonia, hold the paper near the gas without touching the sides of the test tube.
2. Carbon dioxide turns limewater milky by forming calcium carbonate.
3. Chlorine gas bleaches damp litmus paper to a white colour.
4. Hydrogen burns with a characteristic "pop" sound when a burning splint is brought near it.
5. Oxygen causes a glowing splint to reignite and burn more vigorously.
6. Sulphur dioxide decolourises acidified potassium permanganate solution, changing it from purple to colourless.
7. When heated, metal ions emit characteristic colours: lithium crimson red, sodium yellow, potassium lilac, calcium orange-red, barium pale green, copper blue-green.
8. Paper chromatography is a separation technique used to identify and analyse components in a mixture based on their solubility in a solvent.
9. The solvent front is the highest point the solvent reaches on the paper.
10. The baseline, called the origin, is the initial line where the sample is applied.
11. The chromatogram is the visible record of separated components after the experiment.
12. Locating agents are chemicals used to reveal colourless substances. For example, iodine vapour can detect fats and oils.
13. Chromatograms measure the distance each component and the solvent front have travelled from the baseline.

EXERCISE

A. Multiple Choice Questions (MCQs)

Choose the correct option.

i) The gas that turns damp red litmus paper blue is:

(a) carbon dioxide (b) ammonia
(c) chlorine (d) hydrogen

ii) When carbon dioxide gas is passed through limewater, it:

(a) turns red (b) turns milky
(c) turns blue (d) remains clear

iii) The gas that bleaches damp litmus paper is:

(a) hydrogen (b) oxygen
(c) chlorine (d) ammonia

iv) The sound produced when hydrogen gas is tested with a burning splint is:

(a) no sound (b) sizzling
(c) pop sound (d) crackling

v) The flame colour produced by sodium ions (Na^+) is:

(a) lilac (b) intense yellow
(c) crimson red (d) pale green

vi) The main reason for the different flame colours in the flame test is:

(a) different metal cations absorb and emit light at specific wavelengths
(b) the temperature of the flame
(c) the amount of salt used
(d) the colour of the wire loop

vii) The solution that loses its purple colour when sulphur dioxide gas bubbles through it is:

(a) limewater
(b) acidified potassium permanganate
(c) sodium hydroxide solution
(d) dilute hydrochloric acid

viii) The purpose of using a clean Nichrome wire loop in the flame test is to:

(a) increase the flame temperature
(b) avoid contamination and get a pure flame colour
(c) cool the flame
(d) change the colour of the flame

xi) The mobile phase in paper chromatography refers to the:

(a) chromatography paper
(b) solvent that moves up the paper
(c) sample mixture
(d) coloured spots

x) A chromatogram is best described as:

(a) a locating agent
(b) a sample spot on the baseline
(c) the pattern formed by separated substances on the paper
(d) the container used for the chromatography

B. Restricted Response Questions (RRQs)

i) What reagent is used to test for carbon dioxide gas?
ii) How does chlorine gas affect damp litmus paper?
iii) What is the stationary phase in paper chromatography?
iv) What is meant by the term "solvent front"?
v) Calculate the R_f value if a component travels 6cm and the solvent front travels 8cm.

C. Extended Response Questions (ERQs)

I) Describe the reaction between sulphur dioxide gas and acidified potassium permanganate solution. What observation confirms the presence of sulphur dioxide?
ii) Describe the principle of paper chromatography and explain how components of a mixture are separated during the process.
iii) Define the term 'R_f value' and describe how it is calculated and used to identify substances in a mixture.

Answers to Multiple Choice Questions (Units 1-16)

Unit 1:

i) (b) ii) (c) iii) (c) iv) (a) v) (b) vi) (b) vii) (c) viii) (c) ix) (b) x) (c)

Unit 2:

i) (b) ii) (b) iii) (c) iv) (d) v) (b) vi) (b) vii) (b) viii) (c) ix) (c) x) (c)

Unit 3:

i) (c) ii) (b) iii) (b) iv) (c) v) (c) vi) (c) vii) (a) viii) (c) ix) (b) x) (d)

Unit 4:

i) (c) ii) (b) iii) (b) iv) (c) v) (c) vi) (c) vii) (a) viii) (c) ix) (b) x) (d)

Unit 5:

i) (c) ii) (b) iii) (b) iv) (c) v) (a) vi) (c) vii) (b) viii) (a) ix) (b) x) (a)

Unit 6:

i) (b) ii) (b) iii) (a) iv) (a) v) (b) vi) (b) vii) (a) viii) (b) ix) (c) x) (b)

Unit 7:

i) (b) ii) (a) iii) (a) iv) (d) v) (b) vi) (c) vii) (c) viii) (b) ix) (b) x) (b)

Unit 8:

i) (c) ii) (b) iii) (c) iv) (b) v) (d) vi) (c) vii) (b) viii) (a) ix) (c) x) (c)

Unit 9:

i) (c) ii) (c) iii) (c) iv) (c) v) (b) vi) (c) vii) (b) viii) (c) ix) (c) x) (c)

Unit 10:

i) (c) ii) (b) iii) (c) iv) (b) v) (a) vi) (b) vii) (b) viii) (c) ix) (d) x) (d)

Unit 11:

i) (c) ii) (c) iii) (b) iv) (b) v) (a) vi) (a) vii) (b) viii) (b) ix) (c) x) (c)

Unit 12:

i) (c) ii) (c) iii) (c) iv) (b) v) (c) vi) (c) vii) (b) viii) (a) ix) (c) x) (b)

Unit 13:

i) (a) ii) (c) iii) (c) iv) (b) v) (c) vi) (b) vii) (c) viii) (a) ix) (c) x) (c)

Unit 14:

i) (c) ii) (c) iii) (c) iv) (b) v) (c) vi) (c) vii) (c) viii) (a) ix) (d) x) (a)

Unit 15:

i) (c) ii) (c) iii) (b) iv) (c) v) (d) vi) (a) vii) (a) viii) (c) ix) (a) x) (a)

Unit 16:

i) (b) ii) (b) iii) (c) iv) (c) v) (b) vi) (a) vii) (b) viii) (b) ix) (b) x) (c)

References

Philips, J.S., Strozack, V.S., & Wistrom, C. Chemistry: Concept and Application.
Brown, T.L., & Murphy, C.J. Chemistry: The Central Science (13th ed.).
McMurry, J., & Fay, R.C. Chemistry (6th ed.).
Tro, N.J. Chemistry: A Molecular Approach (2nd ed.).
Zumdahl, S.S., & Zumdahl, S.A. Chemistry: An Atom First Approach.
Zumdahl, S.S., & Zumdahl, S.A. Chemistry (8th ed.).
Kotz, J.C., Treichel, P.M., & Townsend, J.R. Chemistry and Chemical Reactivity (8th ed.).
Eubanks, L.P. Chemistry in Context (6th ed.).
Middlecamp, C.H., Keller, S.W., & Hetuel, E.C. Chemistry in Context.
Bailar, C. Chemistry in Your Life (2nd ed.).
Zumdahl, S.S. Chemistry Media – Enhanced Edition.
Masterton, W.L., & Hurley, C.N. Chemistry: Principles and Reactions (6th ed.).
Moore, J., Stanitski, C., & Jurs, P. Chemistry: The Molecular Science (4th ed.).
Gilbert, T.R., Kirss, R., & Foster, N. Chemistry (3rd ed.).
Whitten, K.W., Davis, R.E., Peck, L., & Stanley, G.G. Chemistry (9th ed.).
Gilbert, T.R. Chemistry: The Science in Context (2nd ed.).
Seager, S.L., & Slabaugh, M.R. Chemistry for Today (8th ed.).
Burdge, J. Chemistry (2nd ed.).
Atkins, P., & Jones, L. Chemical Principles (5th ed.).
Cracolice, M.S., & Peters, E.I. Introductory Chemistry.
Tro, N.J. Introductory Chemistry (4th ed.).
Chang, R., & Overby, J. General Chemistry (6th ed.).
Ebbing, D.D., & Gammon, S.D. General Chemistry (9th ed.).
Castellion, M.E., Ballantine, D.S., Hoeger, C.A., & Peterson, R.L. Fundamentals of General, Organic, and Biological Chemistry.
Denniston, K.J. General, Organic, and Biochemistry.
Petrucci, R.H., Herring, F.G., Madura, J.D., & Bissonnette, C. General Chemistry (10th ed.).
Soker, D. General, Organic, and Biological Chemistry (5th ed.).
Raymond, K.W. General, Organic, and Biological Chemistry (4th ed.).
Hein, M., & Arena, J. College Chemistry (13th ed.).
Myers, R.T., Oldham, K.B., & Tocci, S. Chemistry.
Sarquis, M., & Sarquis, J.L. Modern Chemistry.
Decoste, J. Introductory Chemistry (7th ed.).
Jespersen, N.D., Brady, J.E., & Hyslop, A. Chemistry: The Molecular Nature of Matter.
Silberberg, M.S. Principles of General Chemistry (2nd ed.).
Oxtoby, D.W., Gillis, H.P., Campion, A., & Gaither, S. Principles of Modern Chemistry (6th ed.).
Joesten, M.D., Castellion, M.E., & Hogg, J.L. The World of Chemistry (4th ed.).
McMurry, J., & Fay, R.C. General Chemistry (Atoms First) (2nd ed.).
Timberlake, K.C., & Timberlake, W. Chemistry.
Burk, G.R. Introductory Chemistry.
Zeerak, A.J. Chemistry for First Year (Class XI), Balochistan Textbook Board, Quetta.
Additional online sources: Wikipedia, ScienceDirect, Khan Academy, and other educational platforms.

Glossary

Absolute zero: The lowest possible temperature that can be achieved, equal to −273.16 °C or 0 K.

Accuracy: The closeness of a measured value to the true or accepted value.

Acid: A substance that dissociates in water to produce hydrogen ions (H^+).

Activation energy (E_a): The minimum energy required for particles to collide successfully and form products.

Addition reaction: A reaction in which two or more substances combine to form a single product.

Alcohol: An organic compound that contains the hydroxyl (–OH) functional group.

Aldehyde: An organic compound containing the carbonyl (–CHO) group at the end of a carbon chain.

Alkali: A soluble base that produces hydroxide ions (OH^-) in aqueous solution.

Alkali metals: Elements in Group IA (1) of the periodic table.

Alkane: A saturated hydrocarbon with the general formula C_nH_{2n+2}, containing only single bonds.

Alkene: An unsaturated hydrocarbon containing one or more double bonds, with the general formula C_nH_{2n}.

Alkyne: An unsaturated hydrocarbon with one or more triple bonds, with the general formula C_nH_{2n-2}.

Anion: A negatively charged ion formed by gaining electrons.

Anode: The positive electrode where oxidation occurs during electrolysis.

Atom: The smallest particle of an element that retains its chemical identity.

Atomic mass: The weighted average mass of the atoms of an element compared with 1/12th of the mass of a carbon-12 atom.

Atomic number (Z): The number of protons in the nucleus of an atom, determining its identity.

Avogadro's number: The number of particles, 6.022×10^{23}, contained in one mole of a substance.

Base: A substance that can accept hydrogen ions (H^+) or produce hydroxide ions (OH^-) in water.

Boiling point: The temperature at which the vapour pressure of a liquid equals the external pressure.

Bond: The attractive force that holds atoms together in a compound.

Carbohydrate: A biological molecule containing carbon, hydrogen, and oxygen, usually in the ratio $C_n(H_2O)_n$.

Catalyst: A substance that speeds up a chemical reaction without being consumed in the process.

Cathode: The negative electrode where reduction occurs during electrolysis.

Cation: A positively charged ion formed by the loss of electrons.

Covalent bond: A chemical bond formed by sharing electron pairs between atoms.

Crystalline solid: A solid with a regular, repeating arrangement of particles.

Dissociation: The process in which molecules or ionic compounds separate into smaller particles, such as ions.

Distillation: The separation of components of a liquid mixture based on differences in their boiling points.

Double bond: A covalent bond formed by sharing two pairs of electrons between atoms.

Electrolysis: The decomposition of an electrolyte by passing an electric current through it.

Element: A pure substance made up of only one type of atom that cannot be broken down chemically.

Empirical formula: The simplest whole-number ratio of atoms in a compound.

Endothermic reaction: A reaction that absorbs heat energy from its surroundings.

Enzyme: A biological catalyst that speeds up chemical reactions in living organisms.

Equilibrium: A state in which the rate of the forward reaction equals the rate of the reverse reaction.

Exothermic reaction: A reaction that releases heat energy to the surroundings.

Fat: An ester of glycerol with long-chain saturated fatty acids; a type of lipid.

Gas: A state of matter with no fixed shape or volume.

Greenhouse effect: The warming of Earth's atmosphere due to trapped heat by gases like CO_2, CH_4, and water vapour.

Halogens: Reactive non-metals in Group VIIA (17) of the periodic table.

Homogeneous mixture: A mixture with uniform composition and appearance throughout.

Heterogeneous mixture: A mixture in which different components can be visibly distinguished.

Hydrocarbon: An organic compound containing only carbon and hydrogen.

Ion: An atom or group of atoms carrying a positive or negative charge.

Ionic bond: The electrostatic attraction between oppositely charged ions.

Isotope: Atoms of the same element having the same atomic number but different mass numbers.

Kinetic energy: The energy possessed by a moving object due to its motion.

Lattice: A regular, repeating three-dimensional arrangement of particles in a solid.

Mass number (A): The total number of protons and neutrons in the nucleus of an atom.

Metal: An element that is malleable, ductile, lustrous, and a good conductor of heat and electricity.

Mixture: A physical combination of two or more substances that retain their own identities.

Molecular formula: A formula showing the actual number of atoms of each element in one molecule.

Neutralisation: The reaction between an acid and a base to form a salt and water.

Neutron: A neutral subatomic particle found in the nucleus of an atom.

Noble gases: Elements of Group VIIIA (18), which are colourless, odourless, and chemically inert.

Nucleus: The dense, central part of an atom containing protons and neutrons.

Oxidation: The process involving loss of electrons or gain of oxygen.

Oxidising agent: A substance that causes oxidation by accepting electrons from another species.

Periodic table: A systematic arrangement of elements according to their atomic numbers and properties.

pH: It relates to hydrogen ion concentration and is expressed as $pH = -\log[H^+]$.

Precipitate: An insoluble solid formed during a chemical reaction in solution.

Pressure: The force exerted per unit area; measured in pascals (Pa).

Product: A substance formed as a result of a chemical reaction.

Proton: A positively charged subatomic particle found in the nucleus of an atom.

Reactant: A substance that takes part in and undergoes change during a chemical reaction.

Reduction: A process involving the gain of electrons or the loss of oxygen.

Reducing agent: A substance that donates electrons and causes reduction in another substance.

Salt: An ionic compound formed by the neutralisation of an acid and a base.

Saturated solution: A solution that cannot dissolve more solute under the given conditions.

Single bond: A covalent bond involving the sharing of one pair of electrons.

Solubility: The amount of solute that can dissolve in a given quantity of solvent at a specified temperature.

Solution: A homogeneous mixture of solute and solvent.

Solvent: The component of a solution that dissolves the solute, usually in greater amount.

STP: It is abbreviation of standard temperature and pressure. (273.16 K and a pressure of 1 atmosphere.

Symbol: A one- or two-letter abbreviation representing a chemical element.

Temperature: The measure of the average kinetic energy of particles in a substance.

Valence electrons: Electrons present in the outermost shell of an atom, involved in bonding.

Valency: The combining capacity of an element or ion.

Vapour pressure: The pressure exerted by a vapour in equilibrium with its liquid at a given temperature.

Volume: The amount of three-dimensional space occupied by a substance.

About the Author

Mr. Abdullah Jan Zeerak

Mr. Abdullah Jan Zeerak is an Associate Professor of Chemistry at the Government Postgraduate Science College in Quetta. He holds an M.Phil. in Chemistry and has authored several textbooks approved by the Balochistan Textbook Board. He has extensive experience in teaching chemistry at intermediate and undergraduate levels. His interests include chemical education, textbook development, and modern teaching methods. He has participated in many seminars, workshops, and training sessions on curriculum design, teaching, and academic writing.

Mr. Muhammad Shoaib Khan Achakzai

Mr. Shoaib Khan Achakzai is currently employed as a Lecturer in Chemistry at the Government Postgraduate Science College, Quetta. He completed his M.Phil. and PhD in Chemistry. He has published numerous research papers in international journals and has participated in various national conferences. He also developed a chemistry exercise book for intermediate and master's students. Additionally, he has served as a resource person for SLO-based training programmes organised by BACT, with over ten years of teaching experience.