

Q.1. Define Chemistry. What is studied in Chemistry and how does it relate to the properties, composition, and structure of substances?

Ans. Chemistry is the branch of science that studies the properties, composition, and structure of substances. It also explores the physical and chemical changes in matter, as well as the principles and laws that govern these changes. Determining the composition involves finding the percentages of elements and compounds in a sample, while the structure refers to the arrangement of atoms in matter. Both physical and chemical changes can occur through the absorption or release of energy.

Q.2. What are the different branches of Chemistry and their areas of study?

Ans.1. Physical Chemistry:

- (i) Studies how substances behave at atomic and molecular levels.
- (ii) Focuses on the physical laws that govern atoms and molecules.
- (iii) Used to predict and optimize chemical reactions for industrial applications.

2. Inorganic Chemistry:

- (i) Focuses on the synthesis, composition, properties, and structure of elements and compounds that do not contain much carbon.
- (ii) Includes metals, nonmetals, salts, acids, and bases.
- (iii) Applications include fertilizers, medicines, and catalysts.

3. Organic Chemistry:

- (i) Deals with carbon compounds, except simple salts like carbonates and oxides.
- (ii) Studies the structure, properties, and reactions of carbon-containing compounds.
- (iii) Organic compounds are essential for life and found in all living organisms.

4. Environmental Chemistry:

- (i) Investigates the chemical and biochemical phenomena in the environment.
- (ii) Studies the effects of chemicals in air, soil, and water.
- (iii) Helps in understanding pollution and its solutions.

5. Analytical Chemistry:

- (i) Focuses on the analysis of substances.
- (ii) Involves identifying components and determining concentrations.
- (iii) Uses advanced instruments for separation and analysis.

6. Biochemistry:

- (i) Studies the chemical processes in living organisms.
- (ii) Focuses on molecules like proteins, carbohydrates, lipids, and nucleic acids.
- (iii) Provides insights into the molecular structure and function in biological systems.

Interesting Information!

Geothermal heat pump uses a pump to transfer underground water into the buildings during the winter to heat them and in the summer to cool them.

7. Nuclear Chemistry:

- (i) Focuses on nuclear reactions and processes in atomic nuclei.
- (ii) Includes studies of radioactivity and nuclear transformations.
- (iii) Has applications in medicine, agriculture, and industry.

8. Polymer Chemistry:

- (i) Studies the structure, properties, and synthesis of polymers.
- (ii) Polymers are large molecules made of repeating building blocks.
- (iii) Found in materials like proteins, cellulose, and nucleic acids.

9. Geochemistry:

- (i) Studies the chemical composition of Earth's materials, including minerals.
- (ii) Has applications in mineral exploration, environmental monitoring, and medical research.

10. Medicinal Chemistry:

- (i) Focuses on designing and synthesizing drugs beneficial for human health.
- (ii) Involves drug discovery, absorption, and metabolism in the body.

11. Astrochemistry:

- (i) Studies the molecules and ions present in space.
- (ii) Focuses on their reactions and interactions with radiation in the universe.
- (iii) Exercise
- (iv) A lunar mission has recently brought samples from the Moon. The following experiments were then carried out on it. Point out the branch of chemistry these experiments are related to.

Experiment	Branch of Chemistry
Determining its composition	Analytical Chemistry
Studying the physical properties of materials it contains	Physical Chemistry
Carrying out chemical reactions with usual inorganic reagents	Inorganic Chemistry

Q.3. What are the differences between the commonly observed states of matter (solid, liquid, gas, and plasma) and exotic states of matter? Explain with examples.

Ans. **Introduction**

Matter exists in different forms called states, each defined by unique properties such as particle arrangement, intermolecular forces, and behavior under different conditions. Commonly observed states include solids, liquids, gases, and plasma, while exotic states, less encountered in daily life, possess unique and often extreme properties.

Commonly Observed States of Matter

1. Solids

Particle Arrangement: Particles are closely packed in a fixed and orderly structure.



Solid

Properties:

- (i) Definite shape and volume.
- (ii) Strong intermolecular forces.
- (iii) High density and negligible compressibility.

Examples: Metals, rocks, crystalline salts.

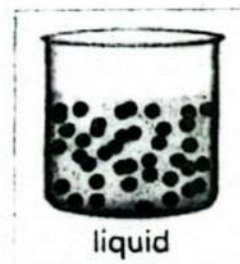
2. Liquids

Particle Arrangement: Particles are close but move randomly.

Properties:

- (i) Takes the shape of the container.
- (ii) Significant intermolecular forces but less than solids.
- (iii) Medium density and low compressibility.

Examples: Water, oil, alcohol.



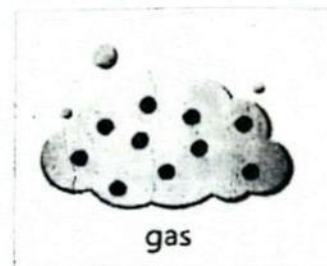
3. Gases

Particle Arrangement: Particles are far apart and move freely.

Properties:

- (i) No fixed shape or volume; expands to fill the container.
- (ii) Weak intermolecular forces.
- (iii) Low density and high compressibility.

Examples: Oxygen, nitrogen, carbon dioxide.



4. Plasma

Particle Arrangement: Particles have extremely high kinetic energy, existing as ions and electrons.

Properties:

- (i) Ionized state of matter.
- (ii) Conducts electricity and reacts to magnetic fields.
- (iii) Found naturally in extreme conditions.

Examples: Lightning, fluorescent lights, stars.

Exotic States of Matter

Exotic states are not commonly observed and usually occur under extreme conditions of temperature or pressure.

1. Bose-Einstein Condensate (BEC)

Description:

Formed at temperatures close to absolute zero. Particles behave as a single quantum entity.

Properties: Extremely low energy, particles are almost motionless.

Applications: Used in research on quantum mechanics.

2. Supercritical Fluids

Description:

A state where matter exhibits properties of both liquids and gases.

Interesting Information!

Many elements are found in nature but some are artificial. Technetium was first element created by scientists in the laboratory.

Properties: Highly compressed; used in specialized chemical reactions.

Examples: Supercritical-carbon dioxide in decaffeination.

3. Liquid Crystals

Description:

A state with properties between liquids and crystalline solids.

Applications: Used in display technology such as LCDs.

4. Graphene

Description:

A single layer of carbon atoms arranged in a hexagonal lattice.

Properties: Strong, flexible, lightweight, and highly conductive.

Applications: Used in electronics and materials science.

5. Dark Matter

Description:

A theoretical form of matter that does not emit, absorb, or reflect light.

Properties: Interacts only through gravity, essential for explaining galaxy formation.

Comparison of Common and Exotic States of Matter

Property	Common States	Exotic States
Occurrence	Observed in everyday life	Found in extreme or theoretical conditions
Examples	Solids, liquids, gases, plasma	BEC, dark matter, supercritical fluids
Behavior	Governed by classical mechanics	Often governed by quantum mechanics
Applications	General use in daily life	Advanced scientific and technological fields

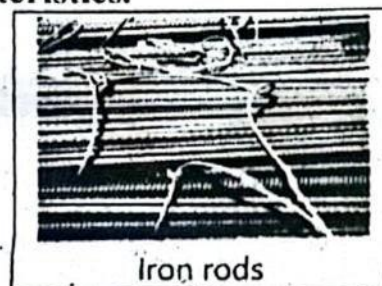
Conclusion

Commonly observed states of matter differ in particle behavior and physical properties, enabling their practical applications in daily life. Exotic states, while rare or theoretical, play a vital role in advancing scientific research and understanding the universe's complexities.

Q.4. What are the differences between elements, compounds, and mixtures? Provide examples for each type of matter and explain their characteristics.

Ans. Introduction

Matter exists in three main forms: elements, compounds, and mixtures. Each type has distinct characteristics based on composition, purity, and the way its components are bonded or mixed. Understanding their differences helps us classify and utilize matter effectively in various applications.



1. Elements

Definition: The simplest form of matter that consists of only one type of atom. Elements cannot be broken down into simpler substances by ordinary chemical reactions.

Characteristics:

- (i) Pure substances.
- (ii) Uniform composition throughout.
- (iii) Exist in solid, liquid, or gaseous states.
- (iv) Can be metals (e.g., copper), non-metals (e.g., oxygen), metalloids (e.g., silicon), or noble gases (e.g., helium).
- (v) Some elements exist as atoms (e.g., sodium), while others exist as molecules (e.g., oxygen gas as O_2).

Examples:

- (i) Solids: Gold (Au), Iron (Fe).
- (ii) Liquids: Mercury (Hg), Bromine (Br).
- (iii) Gases: Nitrogen (N_2), Chlorine (Cl_2).

2. Compounds

Definition: Pure substances formed when two or more different elements chemically combine in a fixed ratio.

Characteristics:

- (i) Uniform composition and properties.
- (ii) Components are bonded through strong chemical bonds (e.g., covalent, ionic).
- (iii) Cannot be separated into individual elements by physical methods; chemical reactions are needed.
- (iv) Classified as organic (e.g., methane) or inorganic (e.g., sodium chloride).

Examples:

Molecular Compounds: Water (H_2O), Ammonia (NH_3).

Ionic Compounds: Sodium chloride (NaCl), Potassium bromide (KBr).

Organic Compounds: Methane (CH_4), Glucose ($C_6H_{12}O_6$).

3. Mixtures

Definition: Formed when two or more elements or compounds are physically combined without any fixed ratio or chemical bonding.

Characteristics:

- (i) Can be separated into components by physical methods (e.g., filtration, distillation).
- (ii) Composition and properties are not uniform unless the



Copper wire



Screw and nuts of zinc

Fig (1.2): Elements



Iron sulphide



Copper chloride

Fig (1.3): Compounds



Rock

mixture is homogeneous.

(iii) Classified as:

- **Homogeneous Mixtures:** Uniform throughout (e.g., saltwater solution).
- **Heterogeneous Mixtures:** Non-uniform, components can be visually distinguished (e.g., rocks).



Examples:

- (i) **Homogeneous Mixtures:** Air, Vinegar.
- (ii) **Heterogeneous Mixtures:** Soil, Salad, Granite.
- (iii) Make comparison between Element, Compound and Mixture.

Comparison Table

Property	Element	Compound	Mixture
Definition	Pure substance of one type of atom.	Pure substance of two or more chemically bonded elements.	Combination of elements/compounds physically mixed.
Composition	Uniform	Uniform	May or may not be uniform.
Bonding	None	Strong chemical bonds	No chemical bonds.
Separation	Not applicable	Requires chemical reaction.	Physical methods.
Examples	Gold, Oxygen	Water, Methane	Air, Soil.

Conclusion

Elements, compounds, and mixtures are fundamental forms of matter, distinguished by their composition, purity, and bonding. Elements and compounds are pure substances with fixed properties, while mixtures are combinations that retain the individual characteristics of their components. This classification is vital in both scientific research and everyday applications.

Q.5. What is allotropy, and what are the allotropic forms of oxygen, carbon, and sulfur? Describe their structures and properties.

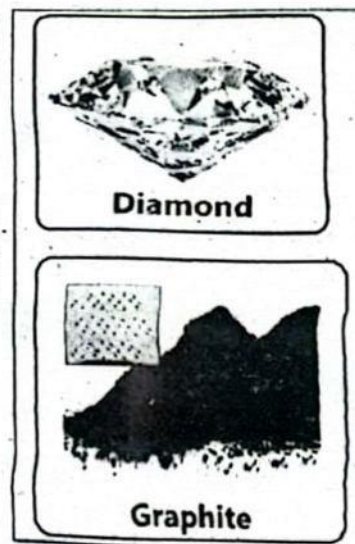
Ans. Definition of Allotropy

Allotropy is the phenomenon in which an element exists in two or more different structural forms. These forms have distinct physical and chemical properties and are called allotropic forms.

Allotropic Forms of Oxygen

1. Oxygen (O₂):

(i) Exists as a diatomic molecule.



(ii) Essential for respiration and combustion.

2. Ozone (O₃):

(i) Composed of three oxygen atoms.

(ii) Acts as a protective layer in the atmosphere by absorbing harmful UV radiation.

Allotropic Forms of Carbon

1. Diamond:

- (i) Giant macromolecular structure.
- (ii) Extremely hard, used in cutting tools.
- (iii) Non-conductor of electricity.

2. Graphite:

- (i) Layered structure with hexagonal rings.
- (ii) Good conductor of electricity.
- (iii) Used as a lubricant and in pencils.

3. Buckminster Fullerene (C₆₀):

- (i) Cage-like spherical structure made of pentagons and hexagons.
- (ii) Soluble in organic solvents.
- (iii) Soft with a low melting point.
- (iv) Cannot conduct electricity.

Allotropic Forms of Sulfur

1. Rhombic Sulfur:

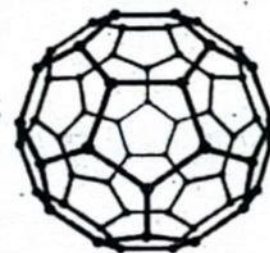
Stable crystalline form under normal conditions.

2. Monoclinic Sulfur:

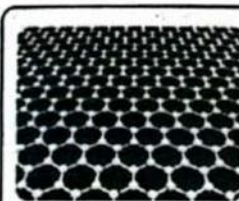
Less stable crystalline form compared to rhombic sulfur.

Q.6. What are the differences between elements, compounds, and mixtures? Present the answer in the properties of tabular form.

Ans.



Buckminsterfullerene



Graphene



Rhombic Sulphur



Monoclinic Sulphur

Fig 1.5 Allotropic

Property	Element	Compound	Mixture
Definition	Simplest form of matter, consisting of the same kind of atoms.	A pure substance formed by a chemical combination of two or more different elements in a fixed ratio.	An impure substance containing a combination of elements or compounds in any ratio.
Purity	Pure substance.	Pure substance.	Impure substance.
Separation	Cannot be broken down into simpler particles by	Can be decomposed into constituent elements by chemical	Components can be separated by physical methods (e.g.,

Property	Element	Compound	Mixture
	ordinary chemical reactions.	reactions.	(filtration, distillation).
Representation	Represented by symbols, e.g., Na (sodium), Ca (calcium).	Represented by chemical formulas, e.g., H ₂ O, NH ₃ .	No fixed representation; components retain their original identity.
Composition	Contains only one type of atom.	Atoms combine in fixed ratios by weight, e.g., in H ₂ O, H:O is 1:8.	Components are mixed in any ratio, e.g., salt and water.
Properties	Possess unique physical and chemical properties.	Properties differ from constituent elements, e.g., water is different from hydrogen and oxygen.	Properties are a combination of the properties of its components.
Bonding	No chemical bonding involved; consists of independent atoms or molecules.	Components are chemically bonded.	Components are not chemically bonded.
Structure	Exists as atoms (e.g., Na) or molecules (e.g., O ₂ , N ₂); noble gases are monoatomic (e.g., He).	Exists as molecules (e.g., HCl) or network structures (e.g., NaCl, SiO ₂).	May be homogeneous (e.g., sugar in water) or heterogeneous (e.g., sand and iron filings).

EXERCISE

1. Which elements are found in pure state on Earth?

Ans. **Elements found in pure state on Earth:**

Gold (Au), Platinum (Pt), Silver (Ag), and Noble gases (e.g., Helium, Neon, Argon).

2. Which elements are present in very small amounts on Earth?

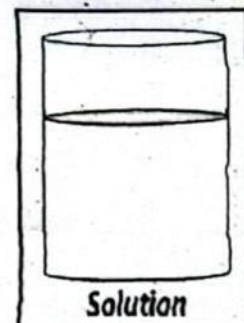
Ans. **Elements present in very small amounts on Earth:**

Scandium (Sc), Technetium (Tc), Astatine (At), and some rare earth elements.

Q.7. What are the differences between a true solution, colloidal solution, and suspension? Explain their properties with examples.

Ans.1. True Solution

Definition: A homogeneous mixture where solute particles are completely dissolved in the solvent.



Example: 203.9g of sugar dissolved in 100g of water at 20°C creates a saturated solution.

Factors Affecting Solubility

1. Nature of Solute and Solvent:

Different substances have different solubilities in the same solvent.

Example: At 20°C, 36g of sodium chloride dissolves in 100g of water, while 203.9g of sugar dissolves.

2. Temperature:

Increasing temperature generally increases solubility for most solids.

3. Molecular Size:

Larger molecules like sugar allow more water molecules to surround and dissolve them, increasing solubility compared to smaller ions like those of sodium chloride.

Q.9. How does temperature affect the solubility of solutes? Explain with examples of different compounds.

Ans. Effect of Temperature on Solubility

General Trend:

For many solutes, solubility increases as temperature increases, but this is not true for all compounds.

Increase in Solubility with Temperature:

Compounds like potassium nitrate (KNO_3), silver nitrate (AgNO_3), and potassium chloride (KCl) have increased solubility with higher temperatures.

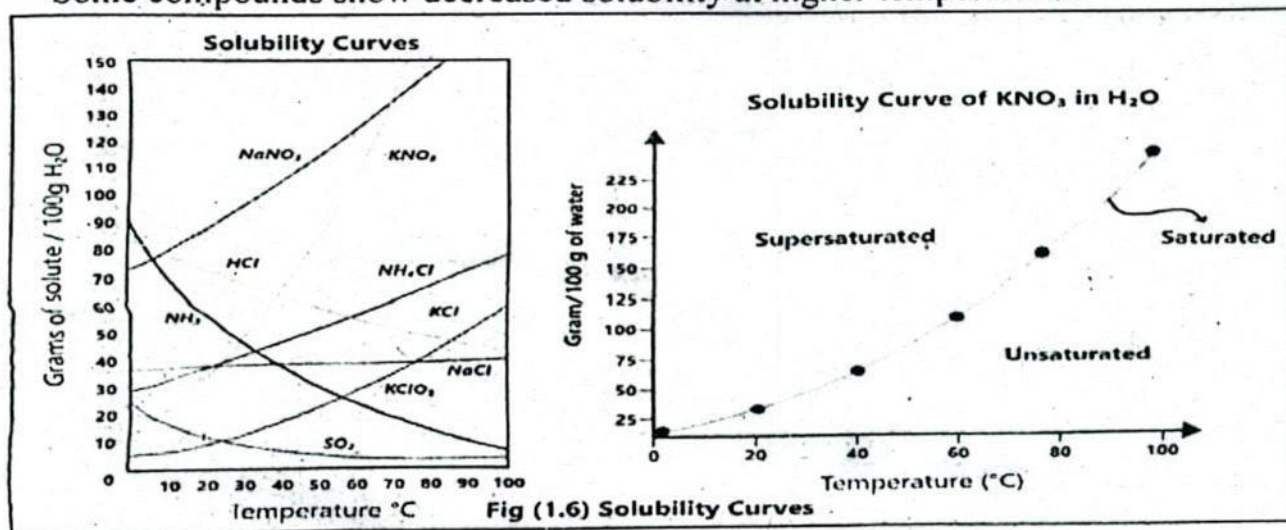
Example: Potassium nitrate's solubility increases in water as the temperature rises.

No Significant Change in Solubility with Temperature Increase:

Example: The solubility of sodium chloride (NaCl) in water does not appreciably change with temperature increases.

Decrease in Solubility with Temperature Increase:

Some compounds show decreased solubility at higher temperatures.



Examples:

(i) Lithium carbonate (Li_2CO_3)

(ii) Calcium chromate (CaCrO_4)

(iii) Calcium hydroxide ($\text{Ca}(\text{OH})_2$)

Example: The solubility of calcium hydroxide decreases as the temperature rises.

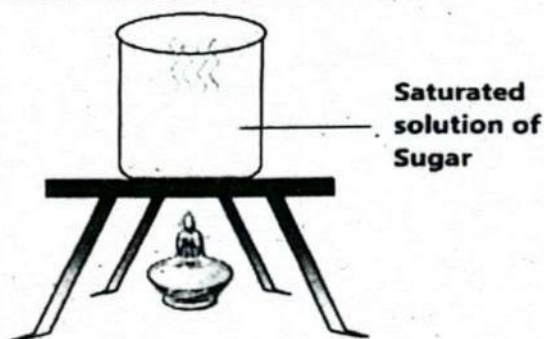
Effect of temperature on solubility of on Gases:

Trend: The solubility of gases in water-generally decreases with an increase in temperature.

Example: Oxygen and carbon dioxide are less soluble in warm water than in cold water.

ACTIVITY

Take 100 g of water in a beaker and prepare saturated solution of sugar at room temperature. Heat the beaker on a spirit lamp. Add a little more sugar in it and stir it. Will this sugar be dissolved in it? You will notice that by heating the solution the quantity of sugar dissolved in water has increased i.e. the solubility of sugar has increased.



Preparing a Saturated Sugar Solution

1. Take 100g of water in a beaker and add sugar to it until no more dissolves. This creates a saturated sugar solution at room temperature.
2. Heat the beaker using a spirit lamp.
3. Add a little more sugar and stir the solution.

Observation:

When the solution is heated, the additional sugar dissolves, showing that the solubility of sugar increases with temperature.

EXERCISE

1. How variation of solubility at different temperatures can be useful for us?

Ans. The variation of solubility with temperature is useful in processes like crystallization, where heating a saturated solution allows more solute to dissolve, and cooling it results in the formation of crystals. This principle is also used in industries like food processing, pharmaceuticals, and chemical manufacturing to control the concentration of solutions or to purify substances.

Interesting Information!

1. The increase in the solubility of solids in liquids with increase in temperature may be used to purify them. Pure solids commonly appear as beautifully shaped crystals.
2. Generally the solubility of gases decreases with increase in temperature. Carbon dioxide gas is more soluble in water at low temperature. Soda water bottles are thus stored in the refrigerator to keep carbon dioxide gas dissolved in water for a longer period of time.



Crystals of Potassium Nitrate

KEY POINTS

1. **What is Chemistry and what does it deal with?**

Ans. Chemistry is the branch of science that deals with the composition of matter, the changes that occur in matter, and the laws that govern these changes.

2. **How is Chemistry divided into different branches?**

Ans. To understand the vast subject of Chemistry, it is divided into several branches. The main branches include physical chemistry, inorganic chemistry, and organic chemistry, among others.

3. **What are the three main states of matter?**

Ans. Matter exists mainly in three states: solid, liquid, and gas. They differ from each other due to the varying distances between the particles they contain.

4. **What is plasma and where is it found?**

Ans. Plasma is regarded as the fourth state of matter, not commonly observed on Earth but is the most prevalent state of matter in the rest of the universe.

5. **What are the intermediate states of matter and give examples?**

Ans. Matter also exists in intermediate states between two principal states, such as between liquid and gas or between liquid and solid. Examples include supercritical fluids and liquid crystals.

6. **How does matter exist in terms of distinct entities?**

Ans. Matter exists in the form of distinct entities called elements, compounds, and mixtures.

7. **What are the differences between elements, compounds, and mixtures?**

Ans. Elements, compounds, and mixtures have distinct properties. They are very different from each other in terms of composition and behavior.

8. **What are the different forms in which mixtures exist?**

Ans. Mixtures typically exist in the forms of solutions, suspensions, and colloidal solutions, each having its own characteristic properties.

9. **What affects the solubility of solutes in a solvent?**

Ans. Different solutes have different solubilities in a particular solvent, meaning the amount of solute that can dissolve varies.

10. **What is solubility and how is it measured?**

Ans. Solubility is the amount of solute that can dissolve in 100g of solvent at a particular temperature.

11. **How does temperature affect the solubility of different compounds?**

Ans. Temperature changes affect the solubility of different compounds in different ways. For many solutes, solubility increases with temperature, but for some, it decreases.

EXERCISE

1. **Tick (✓) the correct answer.**

(i) **Matter is present in neon signs in the state of:**

(a) Supercritical fluid

(b) Plasma

(c) Gas

(d) Liquid crystal

Reason: Neon signs work by using the plasma state of matter, where the gas is ionized, allowing it to conduct electricity and emit light.

(ii) **Hazardous effects of shopping bags are studied in:**

- (a) Geochemistry (b) Inorganic Chemistry
(c) Analytical Chemistry (d) Environmental Chemistry

Reason: Environmental chemistry deals with the study of the effects of chemicals, such as plastic bags, on the environment.

(iii) **The man-made polymer is:**

- (a) Starch (b) Polystyrene
(c) Protein (d) Cellulose

Reason: Polystyrene is a synthetic polymer made from styrene monomers, commonly used in packaging and disposable products.

(iv) **The crystals of which substance has rhombic shape?**

- (a) Brass (b) Sulphur
(c) Graphite (d) Bronze

Reason: Sulphur crystallizes in a rhombic form, which is a characteristic feature of its allotropes.

(v) **Which liquid among the following is a colloidal solution?**

- (a) Milk (b) Slaked lime used for white wash
(c) Vinegar solution (d) Mixture of AgCl in water

Reason: Milk is an emulsion, which is a type of colloidal solution where fat droplets are dispersed in water.

(vi) **Which of the following is a heterogeneous mixture?**

- (a) A solution of calcium hydroxide in water
(b) A solution of potassium nitrate in water
(c) Hot chocolate
(d) Concrete mixture

Reason: Concrete is a heterogeneous mixture because its components (cement, water, sand, etc.) are not uniformly distributed.

(vii) **A state of matter whose properties are between those of liquids and crystalline solids:**

- (a) Liquid crystal (b) Supercritical fluid
(c) Plasma (d) Dark matter

Reason: Liquid crystals have properties of both liquids (flowability) and solids (structure), used in devices like LCD screens.

(viii) **When the tiny visible particles of a substance are dispersed through a medium, the mixture is named as:**

- (a) True solution (b) Colloid
(c) Suspension (d) Saturated solution

Reason: In a suspension, solid particles are dispersed in a liquid or gas but can settle over time, making it different from a true solution.

(ix) **A solution of KClO_3 has a solubility of about 13.2g per 100 cm^3 at 40°C . How its solubility will be affected, if you decrease the temperature?**

- (a) The solubility will increase

- (b) The solubility will decrease
- (c) The solubility will remain the same
- (d) The solubility will first increase with temperature and then it will decrease

Reason: For most solid solutes, solubility decreases as temperature decreases. This is true for $KClO_3$ as well.

(x) **You are studying the rate of hydrolysis of starch under different conditions of temperature. In which branch of chemistry this topic will fall?**

- (a) Organic Chemistry
- (b) Analytical Chemistry
- (c) Biochemistry
- (d) Physical Chemistry

Reason: The study of the rate of reactions and how they change with temperature falls under physical chemistry, which involves the study of chemical reactions and their kinetics.

ANSWERS

(i)	(b)	(ii)	(d)	(iii)	(b)	(iv)	(b)	(v)	(a)
(vi)	(d)	(vii)	(a)	(viii)	(c)	(ix)	(b)	(x)	(d)

Additional MCQ's Question and Answers

1. **Which of the following is a characteristic property of matter?**

- (a) Color
- (b) Density
- (c) Size
- (d) Weight

Reason: Density is a characteristic property of matter because it is unique to a substance and does not depend on the quantity of the substance.

2. **Which branch of chemistry deals with the study of carbon-containing compounds?**

- (a) Organic Chemistry
- (b) Physical Chemistry
- (c) Inorganic Chemistry
- (d) Analytical Chemistry

Reason: Organic chemistry is the branch that focuses on compounds containing carbon, especially hydrocarbons and their derivatives.

3. **Which state of matter has the particles arranged in a regular, repeating pattern?**

- (a) Gas
- (b) Liquid
- (c) Solid
- (d) Plasma

Reason: In solids, the particles are closely packed and arranged in a regular, repeating pattern, giving solids their definite shape and volume.

4. **Which of the following is NOT an example of a colloidal solution?**

- (a) Fog
- (b) Milk
- (c) Salt solution
- (d) Ink

Reason: A salt solution is a true solution, not a colloidal solution, as the solute dissolves completely and the particles are at the molecular level.

5. **What is the fourth state of matter?**

- (a) Gas
- (b) Plasma
- (c) Liquid
- (d) Solid

Reason: Plasma is the fourth state of matter, consisting of ionized gases with free electrons and ions, commonly found in stars and neon signs.

6. Which of the following is a homogeneous mixture?

- (a) Concrete (b) Milk
(c) Salad (d) Oil and water

Reason: Milk is a homogeneous mixture, or colloidal solution, where fat droplets are dispersed in water and uniformly distributed.

7. What happens to the solubility of most solids in water when the temperature increases?

- (a) It decreases (b) It remains the same
(c) It increases (d) It depends on the substance

Reason: For most solid solutes, solubility increases as temperature increases, as it allows for more solute to dissolve.

8. Which of the following is a characteristic of a suspension?

- (a) The particles are uniformly distributed.
(b) The particles are invisible to the naked eye.
(c) The particles settle over time.
(d) The particles do not pass through filter paper.

Reason: In a suspension, the particles are large enough to settle out of the solution over time, unlike in a colloid or solution.

9. Which of these is an example of a heterogeneous mixture?

- (a) Air (b) Salt solution
(c) Sand and water (d) Sugar syrup

Reason: Sand and water form a heterogeneous mixture, as the sand particles do not dissolve and remain separate from the water.

10. Which state of matter has particles that are free to move in all directions?

- (a) Solid (b) Liquid
(c) Gas (d) Plasma

Reason: In the gas state, particles have enough energy to move freely in all directions and spread out to fill the container.

11. In which branch of chemistry is the study of chemical reactions and their rates?

- (a) Organic Chemistry (b) Physical Chemistry
(c) Biochemistry (d) Inorganic Chemistry

Reason: Physical chemistry deals with the study of chemical reactions, their rates, and how temperature, pressure, and concentration affect these reactions.

12. Which of the following is an example of an ionic compound?

- (a) Water (b) Sodium chloride
(c) Carbon dioxide (d) Methane

Reason: Sodium chloride (NaCl) is an ionic compound, formed when sodium donates an electron to chlorine, resulting in oppositely charged ions held together by electrostatic forces.

13. What is formed when a solute dissolves in a solvent to form a homogeneous mixture?

- (a) Suspension (b) Solution
(c) Colloid (d) Emulsion

Reason: A solution is a homogeneous mixture where the solute is dissolved evenly in the solvent, such as sugar in water.

14. **What type of solution can dissolve more solute at a given temperature?**
 (a) Saturated solution (b) Super-saturated solution
 (c) Unsaturated solution (d) Colloidal solution

Reason: An unsaturated solution can dissolve more solute because it contains less solute than the maximum amount it can hold at a given temperature.

15. **What is the effect of increasing pressure on the solubility of gases in liquids?**
 (a) Solubility decreases (b) Solubility increases
 (c) Solubility remains constant (d) There is no relationship

Reason: As pressure increases, the solubility of gases in liquids increases, as more gas molecules are forced into the solution. This is described by Henry's Law.

ANSWERS

1.	(b)	2.	(a)	3.	(c)	4.	(c)	5.	(b)
6.	(b)	7.	(c)	8.	(c)	9.	(c)	10.	(c)
11.	(b)	12.	(b)	13.	(b)	14.	(c)	15.	(b)

Questions for Short Answers

- i. **Why is there a need to divide Chemistry into many branches? Give three reasons.**

Ans. Chemistry is vast and complex, making it easier to study by dividing it into branches.

1. It allows scientists to specialize in specific areas, leading to deeper understanding.
2. Different branches focus on different aspects of matter, helping to address various scientific problems.

- ii. **Reactions may take place due to electrons present outside the nucleus or they may take place inside the nucleus. Which branches of Chemistry cover these two types of reactions?**

Ans. Reactions involving electrons outside the nucleus are studied in **Physical Chemistry** and **Organic Chemistry**.

Reactions occurring inside the nucleus are studied in **Nuclear Chemistry**.

- iii. **What types of problems are solved in analytical chemistry?**

Ans. Analytical chemistry solves problems related to identifying and quantifying substances in a sample, such as determining the concentration of chemicals or detecting impurities.

- iv. **Both graphite and graphene have hexagonal layered structures. What is the difference?**

Ans. Graphite consists of multiple layers of graphene, where the layers are weakly bonded. Graphene is a single layer of carbon atoms arranged in a hexagonal pattern, making it much stronger and more conductive than graphite.

- v. **Why are supercritical fluids important?**

Ans. Supercritical fluids are important because they can dissolve substances like a liquid and diffuse through materials like a gas. This makes them useful in processes like extraction, cleaning, and as solvents in various industries.

- vi. **In which state does matter exist in the Sun?**

Ans. Matter in the Sun primarily exists in the **plasma** state, as the temperatures are extremely high, causing atoms to become ionized.

- vii. **What is the importance of graphene?**
Ans. Graphene is important because it is incredibly strong, lightweight, and conductive. It has a wide range of applications in electronics, energy storage, and material science.
- viii. **Which form of matter do most of the material things in this world belong to?**
Ans. Most material things in the world belong to the solid state of matter.

Additional Short Question and Answers

1. **What is Chemistry?**
Ans. Chemistry is the branch of science that deals with the composition of matter, changes in matter, and the laws that govern these changes.
2. **What are the main branches of Chemistry?**
Ans. The main branches of Chemistry are Physical Chemistry, Inorganic Chemistry, and Organic Chemistry.
2. **What are the three main states of matter?**
Ans. The three main states of matter are solid, liquid, and gas.
3. **What is Plasma?**
Ans. Plasma is considered the fourth state of matter, mostly found in the universe but not commonly on Earth.
4. **What are intermediate states of matter?**
Ans. Intermediate states of matter exist between two principal states, such as between liquid and gas or between liquid and solid, like supercritical fluids and liquid crystals.
5. **What are the basic forms in which matter exists?**
Ans. Matter exists in the form of elements, compounds, and mixtures.
6. **What is a true solution?**
Ans. A true solution is a mixture where the solute particles are completely dissolved in the solvent and cannot be seen by the naked eye.
7. **What is a suspension?**
Ans. A suspension is a mixture in which solute particles do not dissolve but are visible and settle over time.
8. **What is a colloidal solution?**
Ans. A colloidal solution is a mixture where solute particles are larger than in a true solution but smaller than in a suspension, and they do not settle over time.
9. **How does temperature affect the solubility of most solids?**
Ans. The solubility of most solids increases with an increase in temperature.
10. **What is the solubility of a substance?**
Ans. Solubility is the amount of solute that can dissolve in 100g of solvent at a specific temperature.
11. **What happens to the solubility of gases as temperature increases?**
Ans. The solubility of gases decreases with an increase in temperature.
12. **What is the difference between unsaturated and saturated solutions?**
Ans. An unsaturated solution can dissolve more solute, while a saturated solution has dissolved the maximum amount of solute at a given temperature.
13. **Which state of matter has particles that are tightly packed and fixed in place?**
Ans. Solid.

Descriptive Questions

- i. **Mention the name of the branch of Chemistry in which you will study each of the following topics.**
- (a) Rate of a reaction (b) Digestion of food in human body
(c) Properties of plasma (d) Ecosystem
(e) Reactions taking place during fireworks
(f) Measurement of the absorption of wavelength with the help of ultraviolet spectrometer
- Ans. See Chapter
- ii. **What are allotropic forms? Explain the allotropic forms of carbon and sulphur. How does coal differ from diamond?**
- Ans. See Q. No. 1
- iii. **What are supercritical fluids. How are they different from ordinary liquids?**
- Ans. See Q. No. 5
- iv. **Define solubility of a solute. How does the solubility of solutes change with the increase in temperature?**
- Ans. See Q. No. 9
- v. **What types of movements are present in gaseous and liquid molecules?**
- Ans. See Q. No. 6
- vi. **Differentiate between the areas which are studied under inorganic and organic chemistry.**
- Ans. See Q. No.2

Investigative Questions

- i. **Preparation of solutions leads to an important process in chemistry which enables us to purify a compound through crystallization. Describe a process in which potassium nitrate is purified by crystallizing it in water.**
- Ans. To purify potassium nitrate (KNO_3) through crystallization:
1. **Dissolve** impure potassium nitrate in hot water to make a saturated solution.
 2. **Filter** the solution to remove any insoluble impurities.
 3. **Cool** the solution slowly. As it cools, potassium nitrate will crystallize.
 4. **Filter** out the crystals and separate them from the remaining solution (mother liquor).
 5. **Dry** the crystals by allowing them to air dry or using a drying oven.
- This process removes impurities and yields purified potassium nitrate crystals.
- ii. **Graphene is called a miracle material and it is the material of the future. Which of its many properties makes it very useful in electronics?**
- Ans. Graphene is highly useful in electronics due to its **exceptional electrical conductivity**. It allows electrons to move extremely quickly across its surface, making it ideal for high-speed electronic devices. Additionally, its **strength, flexibility, and lightweight nature** make it a promising material for next-generation electronics, such as transistors, sensors, and flexible displays.

electricity through gases at low pressure. These experiments revealed that atoms are not indivisible; they contain smaller particles, including electrons and protons.

4. **Characteristics of Electrons and Protons:**

It was discovered that electrons are negatively charged and are much lighter than protons. A proton is 1836 times heavier than an electron.

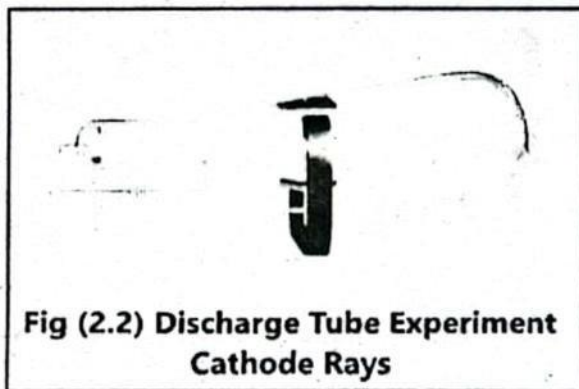
5. **Evidence for Electrons and Protons:**

The presence of negatively charged electrons was confirmed by their deflection towards a positive plate in an electric field. Similarly, protons, which are positively charged, were confirmed by their deflection towards a negative plate.

Q.3. **How were electrons discovered, and what experiments led to their identification as subatomic particles?**

Ans. 1. **Discharge Tube and Cathode Rays:**

A discharge tube consists of a glass tube with metallic electrodes and a vacuum pump. When a high voltage is applied to a gas at low pressure inside the tube, rays are emitted from the cathode, causing the glass surface behind the positive electrode to glow. These rays were named cathode rays.



2. **Thomson's Experiments:**

In 1897, British physicist Joseph John Thomson studied cathode rays by passing them through electric plates with opposite charges. He observed that the cathode rays bent towards the positive plate, indicating they were negatively charged.

3. **Deflection by Magnetic Field:**

Thomson placed magnets on either side of the discharge tube and noticed that the cathode rays were diverted by the magnetic field, further supporting the idea that they were charged particles.

4. **Calculation of Mass-to-Charge Ratio:**

Using his experiments, Thomson calculated the mass-to-charge ratio of cathode rays, which showed that these rays were made of negatively charged material particles. These particles were later named electrons.

5. **Electrons as Subatomic Particles:**

Thomson's work established that electrons are subatomic particles present in all elements, fundamentally changing our understanding of atomic structure.

Conclusion

Through experiments with cathode rays, Joseph John Thomson identified negatively charged particles, which were named electrons. This discovery confirmed that electrons are subatomic particles found in all elements.

Q.4. **How were protons discovered, and what experiments led to their identification as subatomic particles?**

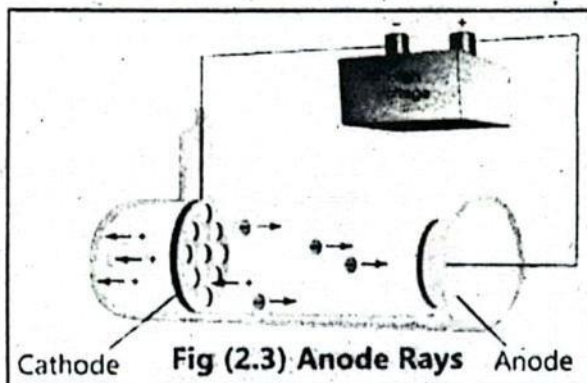
Ans. 1. **Discovery of Canal Rays:**

In 1886, E. Goldstein observed positively charged particles in atoms through

experiments with a gas-discharge tube that had a perforated cathode. The rays produced from the anode moved towards the cathode and were called canal rays or anode rays.

2. Properties of Canal Rays:

The properties of these rays varied depending on the gas used in the discharge tube, revealing the presence of gas ions, including hydrogen ions (H^+). However, Goldstein did not fully understand the significance of his discovery at the time.



3. Rutherford's Experiments (1917):

In 1917, Ernest Rutherford conducted experiments that showed the hydrogen nucleus was present within other atomic nuclei. Rutherford hypothesized that the hydrogen nucleus, which he referred to as a proton, was the fundamental building block of all nuclei.

4. Protons as Fundamental Particles:

Rutherford proposed that the proton was a new fundamental particle that made up the nucleus of atoms, confirming its role as a key subatomic particle.

Conclusion

E. Goldstein's discovery of canal rays led to the identification of positively charged particles, and Rutherford's experiments further confirmed the existence of protons as the fundamental building blocks of atomic nuclei.

Q.5. How were neutrons discovered, and what is the structure of an atom based on the discovery of fundamental particles?

Ans. 1. Discovery of Neutron (1933):

In 1933, the neutron was discovered, a subatomic particle that carries no charge. Its mass is nearly the same as that of a proton, making it an essential component of atomic nuclei.

2. Fundamental Particles:

The three fundamental particles—electron, proton, and neutron—are present in all atoms (except ordinary hydrogen, which lacks a neutron). Despite their different behaviors, these particles make up the structure of all atoms.

3. Variation in Number of Particles:

The number of electrons, protons, and neutrons varies in different atoms, which leads to the diversity of elements and their distinct properties.

4. Exploration of Atomic Structure:

After the discovery of these subatomic particles, a critical question arose: how are these particles arranged in the small space of the atom? This led to the exploration of the atomic structure.

Conclusion

The discovery of the neutron, along with the electron and proton, led to the understanding that all atoms are made of these fundamental particles. The

arrangement of these particles within the atom became a key question in understanding atomic structure.

Q.6. How did Lord Rutherford discover the nucleus, and what are the properties of the particles in the atom?

Ans. 1. Rutherford's Experiment (1911):

In 1911, Lord Rutherford conducted a groundbreaking experiment where he directed a stream of special particles at a very thin gold foil. From this experiment, he concluded that the atom consists of two main parts: a tiny central portion called the nucleus and a much larger surrounding region known as the extra-nuclear portion.

2. Discovery of the Nucleus:

Rutherford's findings showed that the nucleus is extremely small compared to the entire atom. It was also discovered that almost all the mass of an atom is concentrated in the nucleus, as it contains the heavy particles, protons and neutrons.

3. Nuclear Force:

Inside the nucleus, protons and neutrons are held together by a strong nuclear force, which is responsible for the stability of the nucleus.

4. Properties of Subatomic Particles:

The following are the properties of the fundamental particles in an atom:

Particle	Charge	Mass
Electron	$-1.6022 \times 10^{-19} \text{C}$	$9.109 \times 10^{-31} \text{Kg}$
Proton	$+1.6022 \times 10^{-19} \text{C}$	$1.673 \times 10^{-27} \text{Kg}$
Neutron	0.0	$1.675 \times 10^{-27} \text{Kg}$

Conclusion

Lord Rutherford's experiment led to the discovery of the atomic nucleus, where most of the atom's mass is concentrated, and he identified the strong nuclear force that holds protons and neutrons together in the nucleus. The properties of the electron, proton, and neutron were also defined, providing a clearer understanding of atomic structure.

Interesting Information!

Although the nucleus is less than one hundred-Thousandth (1/100,000) of the size of the atom, it contains more than 99.9% of the mass of the atom.

Q.7. What is Bohr's atomic model, and how does it explain the arrangement of electrons in an atom?

Ans.1. Bohr's Atomic Model (1913):

Niels Bohr proposed that electrons revolve around the nucleus in specific paths called orbits or shells. Each orbit has a fixed energy, and electrons can only exist in these discrete energy levels.

2. Ground State and Energy Levels:

The electron in the orbit closest to the nucleus has the minimum energy and is in the ground state. Orbits further from the nucleus have progressively greater energy.

3. **Fixed Energy of Electrons in Orbits:**

Each shell has a fixed energy, and the electron cannot occupy any space between the orbits. The shells closest to the nucleus (like the first shell, K shell) have lower energy than those further away (like the second, L shell).

4. **Sub-shells and Orbitals:**

Each shell is subdivided into sub-shells (orbitals). The number of sub-shells in a shell is equal to the shell's number (n):

n = 1 (K shell) has one sub-shell (s).

n = 2 (L shell) has two sub-shells (s and p).

n = 3 (M shell) has three sub-shells (s, p, and d).

n = 4 (N shell) has four sub-shells (s, p, d, and f).

5. **Modern Understanding of Electrons:**

In the modern approach, electrons are described as charged clouds, and their exact location around the nucleus cannot be predicted with 100% certainty. Instead, the probability of finding an electron at a certain distance from the nucleus is considered.

6. **Maximum Number of Electrons in Shells:**

The maximum number of electrons that can be accommodated in each shell is given by the formula $2n^2$, where n is the shell number:

K shell (n = 1) can hold 2 electrons.

L shell (n = 2) can hold 8 electrons.

M shell (n = 3) can hold 18 electrons.

N shell (n = 4) can hold 32 electrons.

7. **Electrons in Sub-shells:**

Each sub-shell can hold a fixed number of electrons:

s sub-shell: 2 electrons.

p sub-shell: 6 electrons.

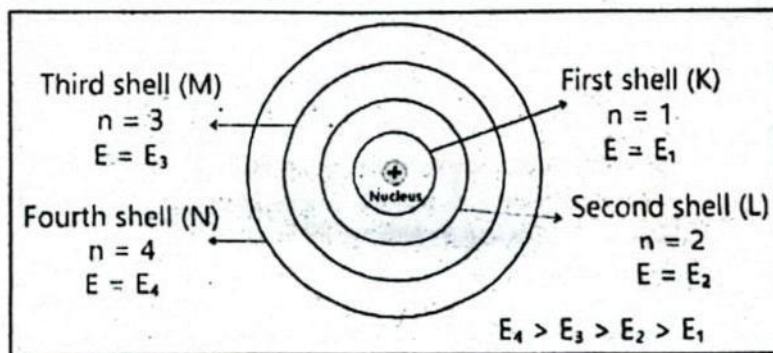
d sub-shell: 10 electrons.

f sub-shell: 14 electrons.

Conclusion

Bohr's atomic model explains that electrons exist in specific orbits around the nucleus, with each orbit corresponding to a distinct energy level.

These orbits are further divided into sub-shells, and the number of electrons that can fit in each shell is determined by the formula $2n^2$. Modern atomic theory also describes electrons as existing in probable regions around the nucleus rather than fixed paths.



Interesting Information!

The size of an atom is so small that it is not possible to see it with naked eye. However, a transition electron microscope can be used to see atoms.

Interesting Information!

The size of an atom is so small that it is not possible to see it with naked eye. However, a transition electron microscope can be used to see atoms.

Q.8. What are atomic number and mass number, and how are they used to identify elements and determine the number of neutrons in an atom?

Ans. 1. Fundamental Particles of Matter:

Electrons, protons, and neutrons are the fundamental particles of all types of matter. Each atom of an element contains these particles, but the number of these particles varies between different elements.

2. Atomic Number (Z):

The atomic number of an element is the number of protons in the nucleus of an atom. It is unique to each element and determines the element's identity. Since atoms are electrically neutral, the number of electrons in an atom is equal to the number of protons. The atomic number is represented by "Z."

3. Mass Number (A):

The mass number, or nucleon number, is the total number of protons and neutrons in an atom. It is represented by "A." The mass of electrons is negligible, so it is not included in the mass number. The mass number helps to determine the overall mass of an atom.

4. Example with Oxygen:

For an oxygen atom, the atomic number (Z) is 8 because it has 8 protons, and the mass number (A) is 16 because it has a total of 8 protons and 8 neutrons. Oxygen can be symbolized as ${}^16_8\text{O}$.

5. Identifying Elements with Atomic and Mass Numbers:

The atomic number and mass number are used to identify elements. For example, a carbon atom, symbolized as ${}^{12}_6\text{C}$, has 6 protons and 6 neutrons.

6. Calculating Neutrons (N):

The number of neutrons in an atom can be calculated using the formula:

$$N = A - Z$$

For example, in a sulfur atom with a mass number of 32 and an atomic number of 16, the number of neutrons is:

$$N = 32 - 16 = 16$$

Conclusion

The atomic number (Z) is the number of protons in an atom and identifies the element. The mass number (A) is the total number of protons and neutrons. The number of neutrons can be found by subtracting the atomic number from the mass number.

Interesting Information!

Changing the number of electrons of an element forms ions, while changing the number of neutrons of an element forms isotopes.

Example:

Calculate the number of neutrons, protons and electrons in barium $^{137}_{56}\text{Ba}$

Solution: Each barium atom will have 56 protons and 56 electrons. The number of neutrons in barium will be calculated as follows:

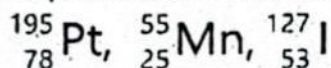
$$N = A - Z$$

$$N = 137 - 56 = 81$$

So each $^{137}_{56}\text{Ba}$ atom will have 81 neutrons, 56 protons and 56 electrons.

EXERCISE

Calculate the number of neutrons, protons and electrons in the following atoms.

**Sol:1. Platinum (Pt):**

(i) **Atomic Number (Z)** = 78 (This is the number of protons)

(ii) **Mass Number (A)** = 195

Number of Protons = 78

Number of Electrons = 78 (Since the atom is neutral, the number of electrons is equal to the number of protons)

Number of Neutrons = Mass Number - Atomic Number = $195 - 78 = 117$
Neutrons

2. Manganese (Mn):

(i) **Atomic Number (Z)** = 25 (This is the number of protons)

(ii) **Mass Number (A)** = 127

Number of Protons = 25

Number of Electrons = 25 (Since the atom is neutral, the number of electrons is equal to the number of protons)

Number of Neutrons = Mass Number - Atomic Number = $127 - 25 = 102$
Neutrons

Q.9. What are isotopes, and how do their mass numbers differ? Provide examples with isotopes of carbon and hydrogen.

Ans.1. Isotopes:

Isotopes are atoms of the same element that have the same atomic number but different mass numbers. This happens because they have the same number of protons but different numbers of neutrons in their nuclei.

2. Isotopes of Carbon:

Carbon has three isotopes, which differ in the number of neutrons:

Carbon-12 (^{12}C): 6 protons and 6 neutrons.

Carbon-13 (^{13}C): 6 protons and 7 neutrons.

Carbon-14 (^{14}C): 6 protons and 8 neutrons.

These isotopes have different mass numbers: 12, 13, and 14, respectively.

3. Isotopes of Hydrogen:

Hydrogen has three isotopes:

Hydrogen (${}^1\text{H}$): 1 proton and no neutrons (the most common isotope).

Deuterium (${}^2\text{H}$ or **D):** 1 proton and 1 neutron.

Tritium (${}^3\text{H}$ or **T):** 1 proton and 2 neutrons.

4. **Physical vs. Chemical Properties:**

The chemical properties of isotopes are nearly identical because they have the same number of electrons, which determine their chemical behavior. However, their physical

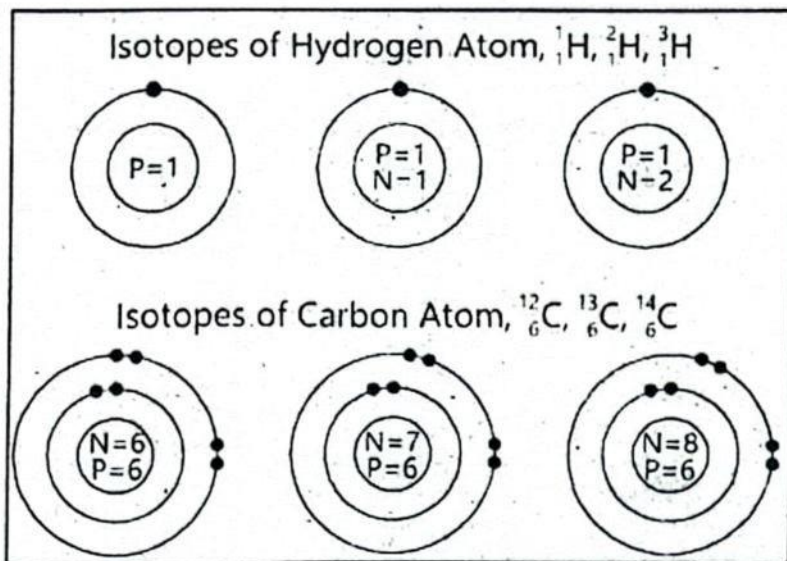
properties, such as mass and stability, differ due to the different numbers of neutrons.

For example:

- (i) Deuterium has twice the mass of Hydrogen.
- (ii) Tritium has three times the mass of Hydrogen.

5. **Mass Difference in Isotopes:**

Each isotope of an element has a different mass number, which reflects the number of protons and neutrons in the nucleus. However, the chemical behavior remains very similar due to the same number of electrons.



EXERCISE

1. Why isotopes of an element show same chemical properties while their physical properties are different?

Ans. Isotopes of an element show the same chemical properties because they have the same number of electrons, which determines how they interact with other elements. However, their physical properties differ due to variations in their mass and the number of neutrons, which affect properties like density, melting point, and stability.

Q.10. What are radioactive isotopes and how do they undergo radioactive decay?

Ans. 1. Radioactive Isotopes:

- (i) Isotopes are atoms of the same element that have the same number of protons but a different number of neutrons. These isotopes can have different physical properties.
- (ii) Some isotopes are unstable, meaning their nuclei do not have enough binding energy to hold together, causing them to release excess energy in the form of radiation.
- (iii) This release of energy is called **radioactivity**, and the isotopes that emit radiation are known as **radioactive isotopes**.

2. **Radioactive Decay:**

During radioactive decay, the unstable isotope undergoes transformation into a new element by emitting radiation. The process continues until a stable isotope is formed.

4.

Energy Production

Radioactive isotopes are used in nuclear reactors to generate cheap nuclear power.

Interesting Information!

Gallium has many interesting properties. Its melting point is below body temperature so it is liquid at room temperature. It does not evaporate.

EXERCISE

1. **Why does a radioactive isotope emit radiation?**

Ans. A radioactive isotope emits radiation because its nucleus is unstable. To achieve stability, it releases excess energy in the form of radiation through a process called radioactive decay.

2. **Give an example of a radioactive isotope that disintegrates to give a stable atom.**

Ans. Uranium -238 (${}_{92}^{238}\text{U}$) disintegrates to form Thorium -234 (${}_{90}^{234}\text{Th}$) and eventually transforms into Lead -206 (${}_{82}^{206}\text{Pb}$), a stable atom.

Q.12. **What happens when a sodium atom is hit by ionizing radiation emitted by a radioactive source?**

Ans. **Process of Ionization**

When a sodium atom is struck by ionizing radiation, such as the radiation emitted by radium-226, the energy of the radiation may be sufficient to remove an electron from the atom's outermost orbit.

Formation of a Cation

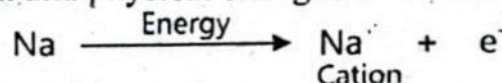
The loss of an electron results in the sodium atom becoming positively charged, forming a cation (Na^+).

Cause of Ionization

The ionization occurs because the radiation provides enough energy to overcome the binding energy of the electron, allowing it to be ejected from the atom.

Significance

This process demonstrates how ionizing radiation can alter the electronic structure of atoms, leading to chemical and physical changes in the affected material.

**Interesting Information!**

The same elements occur everywhere in the universe. Matter on Mars or in any other planet consists of the same elements found on Earth.

Q.13. **How is the relative atomic mass of an element determined, and why is it important in chemistry?**

Ans. **Definition of Relative Atomic Mass**

The relative atomic mass of an element is the mass of an atom of that element compared to 1/12th the mass of an atom of the carbon -12 isotope. The carbon -12 isotope is taken as the standard reference, with its mass defined as exactly 12 atomic mass units (amu).

Importance of Relative Atomic Mass

Knowing the relative atomic masses allows chemists to determine the ratios of masses in which elements combine during chemical reactions. This knowledge is crucial for accurately measuring and mixing reactants in proper proportions.

Measurement Unit

The relative atomic mass is expressed in atomic mass units (amu), where:

$$1 \text{ amu} = 1.67377 \times 10^{-27} \text{ kg.}$$

Examples

Hydrogen - 1: The mass of one atom of hydrogen -1 on the atomic mass scale is 1.007 amu.

Sulfur - 32: The mass of one atom of sulfur -32 on the same scale is 31.972 amu.

Significance

This scale provides a uniform system for comparing atomic masses, facilitating chemical calculations and reactions across various elements.

EXERCISE

1. How would you compare the masses of the atoms of Mg and Cl?

Ans. To compare the masses of magnesium (Mg) and chlorine (Cl) atoms, use their relative atomic masses:

- **Magnesium (Mg):** Relative atomic mass = 24.305 amu
- **Chlorine (Cl):** Relative atomic mass = 35.45 amu

Thus, a chlorine atom is approximately 1.46 times heavier than a magnesium atom.

Q.14. How can the relative atomic mass of krypton (Kr) be calculated from its isotopic masses and abundances?

Ans. **Formula for Calculation**

The relative atomic mass (RAM) of an element is calculated using the following formula:

$$\text{RAM} = \frac{(m_1 \times p_1) + (m_2 \times p_2) + \dots + (m_n \times p_n)}{100}$$

Where:

- m_1, m_2, \dots, m_n are the relative isotopic masses of the isotopes.
- p_1, p_2, \dots, p_n are the isotopic abundances of the isotopes.

Given Isotopic Data for Krypton (Kr)

Krypton has five isotopes with the following relative isotopic masses and abundances:

Relative Isotopic Mass	Isotopic Abundance
80	2.0%
82	12.0%
83	12.0%
84	57.0%
85	17.0%

Step-by-Step Calculation

To calculate the relative atomic mass (RAM) of krypton, we substitute the given values into the formula:

$$\text{RAM} = \frac{(80 \times 2.0) + (82 \times 12.0) + (83 \times 12.0) + (84 \times 57.0) + (85 \times 17.0)}{100}$$

$$\text{RAM} = \frac{160 + 984 + 996 + 4788 + 1445}{100} = \frac{7373}{100} = 73.73$$

Conclusion

The relative atomic mass of krypton (Kr) is **73.73 amu**.

Sample Example 1:

Relative atomic mass of krypton

$$= \frac{80 \times 2.0 + 82 \times 12.0 + 83 \times 12.0 + 84 \times 57.0 + 85 \times 17.0}{100}$$

$$= 83.7$$

Sample Example 2:

Calculate the relative atomic mass of light isotope of chlorine when its relative atomic mass is taken as 35.45.

Relative isotopic abundance of Cl-37 = 24.23%

Relative isotopic abundance of light isotope of chlorine = 75.77%

Relative atomic mass of chlorine = 35.45

Relative atomic mass of chlorine = $\frac{\text{Cl} \times 75.77 + 37 \times 24.23}{100}$

$$35.45 = \frac{\text{Cl} \times 75.77 + 37 \times 24.23}{100}$$

$$3545 = \text{Cl} \times 75.77 + 37 \times 24.23$$

$$3545 - 896.51 = \text{Cl} \times 75.77$$

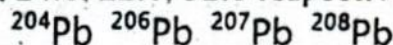
$$2648.49 / 75.77 = \text{Cl}$$

$$34.95 = \text{Cl}$$

Relative atomic mass of light isotope of chlorine is 34.95.

EXERCISE

1. Calculate the relative atomic mass of Lead (Pb). Isotopic abundances of isotopes are 2.0, 24.0, 22.0, 52.0 respectively.



To calculate the relative atomic mass of lead (Pb) using its isotopes ${}^{204}\text{Pb}$, ${}^{206}\text{Pb}$, ${}^{207}\text{Pb}$, and ${}^{208}\text{Pb}$ with the given isotopic abundances, follow these steps:

Given Data:

- Isotopes of Pb: ${}^{204}\text{Pb}$, ${}^{206}\text{Pb}$, ${}^{207}\text{Pb}$, ${}^{208}\text{Pb}$
- Isotopic Masses: 204, 206, 207, 208
- Isotopic Abundances: 2.0%, 24.0%, 22.0%, 52.0%

Formula for Relative Atomic Mass:

$$\text{RAM} = \frac{(m_1 \times p_1) + (m_2 \times p_2) + (m_3 \times p_3) + (m_4 \times p_4)}{100}$$

Where:

- $m_1 = 204$, $m_2 = 206$, $m_3 = 207$, $m_4 = 208$ are the isotopic masses.
- $p_1 = 2.0\%$, $p_2 = 24.0\%$, $p_3 = 22.0\%$, $p_4 = 52.0\%$ are the isotopic abundances.

Step-by-Step Calculation:

$$\text{RAM} = \frac{(204 \times 2.0) + (206 \times 24.0) + (207 \times 22.0) + (208 \times 52.0)}{100}$$

$$\text{RAM} = \frac{408 + 4944 + 4554 + 10816}{100}$$

$$\text{RAM} = \frac{20322}{100} = 203.22$$

Conclusion:

The relative atomic mass of lead (Pb) is **203.22 amu**.

KEY POINTS

1. Who provided the evidence for the existence of atoms, and how was this evidence gathered?

Ans. John Dalton, an English chemist, provided evidence for the existence of atoms through his work in the early 19th century. His experiments on the behavior of gases led him to propose that matter is made up of indivisible particles called atoms, each with a unique weight.

2. What did discharge tube experiments reveal about atoms?

Ans. Discharge tube experiments showed that atoms are not the smallest particles of matter. Instead, atoms are composed of smaller particles, namely electrons and protons. Neutrons were later discovered as another subatomic particle.

3. What was discovered about the subatomic particles present in all elements?

Ans. It was discovered that all elements contain electrons, protons, and neutrons, regardless of how differently the elements behave chemically. The number of these particles, however, varies across different elements.

4. What did Lord Rutherford discover about the structure of the atom?

Ans. Lord Rutherford discovered that all atoms have a central part called the nucleus, which contains protons and neutrons. Electrons revolve around this nucleus.

5. Why is an atom electrically neutral?

Ans. An atom is electrically neutral because it contains an equal number of electrons and protons. The negative charge of electrons balances the positive charge of protons.

6. What does Bohr's atomic model explain about electron behavior?

Ans. Bohr's atomic model explains that electrons revolve around the nucleus in specific, fixed orbits or shells. An electron's energy remains constant as long as it stays in one orbit. Electrons located farther from the nucleus have higher energy.

7. **What is the atomic number of an element?**
Ans. The atomic number of an element is the number of protons present in the nucleus of an atom of that element. It defines the element and determines its position in the periodic table.
8. **What is the mass number of an element?**
Ans. The mass number of an element is the total number of protons and neutrons in the nucleus of an atom of that element.
9. **What are isotopes?**
Ans. Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons. This results in different mass numbers for each isotope.
10. **How do isotopes of an element differ from each other?**
Ans. Isotopes of an element have the same chemical properties because they have the same number of electrons. However, they differ in their physical properties, mainly due to differences in mass.
11. **What are the two types of isotopes, and how are they significant?**
Ans. Isotopes can be stable or radioactive. Radioactive isotopes have unstable nuclei and release radiation. They are used in various applications, particularly in medicine for diagnosis and treatment.
12. **What is the significance of radioactive isotopes?**
Ans. Radioactive isotopes have unstable nuclei and emit radiation as they decay. These isotopes have useful applications in fields like medicine, where they are used in imaging, cancer treatment, and sterilization.
13. **How is the relative atomic mass of an element calculated?**
Ans. The relative atomic mass of an element is calculated by considering the relative isotopic masses of the element's isotopes and their isotopic abundances. This can be done using the weighted average formula.

EXERCISE

1. **Tick (✓) the correct answer.**
- (i) **How many electrons can be accommodated at the most in the third shell of the elements?**
- | | |
|--------|--------|
| (a) 8 | (b) 18 |
| (c) 10 | (d) 32 |
- Reason:** The third shell ($n = 3$) can hold a maximum of 18 electrons, calculated using the formula $2n^2$, where $n = 3$, so $2(3)^2 = 18$.
- (ii) **What information was obtained from discharge tube experiments?**
- (a) Structure of atom was discovered.
 - (b) Neutrons and protons were discovered.
 - (c) Electrons and protons were discovered.
 - (d) Presence of nucleus in an atom was discovered.

Reason: Discharge tube experiments led to the discovery of electrons and protons by

observing the behavior of cathode rays and positive rays.

(iii) **Why have isotopes not been shown in the periodic table?**

- (a) Periodic table cannot accommodate a large number of isotopes of different elements.
- (b) Some of the isotopes are unstable and they give rise to different elements
- (c) All the isotopes have same atomic number; so there is no need to give them separate places.
- (d) Isotopes do not show periodic behavior.

Reason: Isotopes of an element have the same atomic number (number of protons) but different mass numbers (number of neutrons), so they are represented by a single position on the periodic table.

(iv) **Which particle is present in different number in the isotopes?**

- (a) Electron
- (b) Neutron
- (c) Proton
- (d) Both neutron and electron

Reason: Isotopes of an element differ in the number of neutrons they contain, but they have the same number of protons and electrons.

(v) **In which isotope of oxygen there are the equal number of protons, electrons and neutrons?**

- (a) ^{17}O
- (b) ^{16}O
- (c) ^{18}O
- (d) None of these

Reason: Oxygen -16 (^{16}O) has 8 protons, 8 electrons, and 8 neutrons, making the number of protons, electrons, and neutrons equal.

(vi) **What will be the relative atomic mass of nitrogen given the abundances of its two isotopes, ^{14}N and ^{15}N are 99.64 and 0.35 respectively.**

- (a) 14.0210
- (b) 14.0021
- (c) 14.2100
- (d) 14.1200

Reason: The relative atomic mass is calculated as the weighted average of the isotopes:

$$\text{Relative atomic mass} = (14 \times 0.9964) + (15 \times 0.0035) = 14.0021$$

(vii) **How is radiocarbon dating useful for archeologists?**

- (a) It helps determine the age of organic matter.
- (b) It helps determine the composition of matter.
- (c) It helps determine the usefulness of matter.
- (d) It helps determine whether the matter is radioactive or not.

Reason: Radiocarbon dating measures the amount of carbon -14 in organic material to estimate its age, which is particularly useful in archaeology.

(viii) **What does keep the particles present in the nucleus intact?**

- (a) Particles are held together by strong nuclear force.
- (b) Particles are held together by weak nuclear force.

- (c) Particles are held together by electrostatic force.
- (d) Particles are held together by dipolar force.

Reason: The strong nuclear force is the force that holds protons and neutrons together in the nucleus, overcoming the electrostatic repulsion between positively charged protons.

(ix) **How do electrons keep themselves away from the oppositely charged nucleus?**

- (a) By keeping themselves stationary
- (b) By revolving around the nucleus
- (c) Due to their wave-like nature
- (d) A magnetic field around the nucleus keeps them away

Reason: Electrons do not spiral into the nucleus because they exhibit wave-like behavior, as described by the quantum mechanical model, which keeps them in orbitals at specific distances from the nucleus.

(x) **Rubidium consists of two isotopes ^{85}Rb and ^{87}Rb . The percent abundance of the light isotope is 72.2% What is the percent abundance of the heavier isotope? Its atomic mass is 85.47**

- (a) 15%
- (b) 27.8%
- (c) 37%
- (d) 72%

Reason: Since the total abundance must sum to 100%, the percent abundance of the heavier isotope ^{87}Rb is $100 - 72.2 = 27.8\%$.

ANSWERS

(i)	(b)	(ii)	(c)	(iii)	(c)	(iv)	(b)	(v)	(b)
(vi)	(b)	(vii)	(a)	(viii)	(a)	(ix)	(c)	(x)	(b)

Additional MCQ's Question and Answers

1. **What is the atomic number of an element?**

- (a) Number of electrons
- (b) Number of protons
- (c) Number of neutrons
- (d) Total number of protons and neutrons

Reason: The atomic number of an element is equal to the number of protons in the nucleus of its atoms. It determines the identity of the element.

2. **Which subatomic particle has a positive charge?**

- (a) Electron
- (b) Neutron
- (c) Proton
- (d) All of the above

Reason: Protons are positively charged particles found in the nucleus of the atom, while electrons are negatively charged and neutrons are neutral.

3. **Which part of the atom was discovered by Rutherford?**

- (a) Electrons
- (b) Neutrons
- (c) Nucleus
- (d) Protons

Reason: Rutherford's gold foil experiment led to the discovery of the atomic nucleus, which is the dense, positively charged center of the atom.

4. **What is the mass number of an element?**

- (a) The number of protons (b) The number of neutrons
(c) The number of protons and neutrons (d) The number of electrons

Reason: The mass number is the sum of the number of protons and neutrons in an atom's nucleus.

5. **Which model of the atom proposed that electrons revolve in fixed orbits?**

- (a) Dalton's model (b) Thomson's model
(c) Bohr's model (d) Rutherford's model

Reason: Bohr's model of the atom proposed that electrons revolve in specific orbits around the nucleus, each corresponding to a specific energy level.

6. **Which of these particles is not found in the nucleus of an atom?**

- (a) Proton (b) Neutron
(c) Electron (d) Both a and b

Reason: Electrons are found outside the nucleus in orbitals, while protons and neutrons are located inside the nucleus.

7. **Which of the following elements contains an atom with 6 protons and 6 neutrons?**

- (a) Carbon(b) Oxygen
(c) Nitrogen(d) Hydrogen

Reason: Carbon-12 isotope has 6 protons and 6 neutrons, making it the most common isotope of carbon.

8. **What does the atomic mass unit (amu) represent?**

- (a) The mass of an electron (b) The mass of a proton
(c) One-twelfth of the mass of a carbon-12 atom
(d) The mass of a neutron

Reason: The atomic mass unit (amu) is defined as one twelfth the mass of a carbon-12 atom, which is used to express atomic and molecular weights.

9. **Which of these is not a property of isotopes?**

- (a) Same number of protons (b) Same chemical properties
(c) Same number of neutrons (d) Different mass numbers

Reason: Isotopes of an element have the same number of protons but a different number of neutrons, leading to different mass numbers.

10. **What type of isotopes are used in radioactive dating?**

- (a) Stable isotopes (b) Radioactive isotopes
(c) Both stable and radioactive isotopes
(d) Non-radioactive isotopes

Reason: Radioactive isotopes, such as carbon-14, are used in radioactive dating because they decay at a predictable rate, allowing scientists to estimate the age of materials.

11. **What is the relationship between the number of protons and the atomic number?**

- (a) They are equal (b) Protons are half of the atomic number
(c) Protons are double the atomic number
(d) They are inversely related

Reason: The atomic number is the number of protons in an atom's nucleus, so they are equal by definition.

12. **What happens when an atom loses an electron?**

- (a) It becomes a negatively charged ion

- (b) It becomes a positively charged ion
 (c) It becomes neutral (d) It becomes a molecule

Reason: When an atom loses an electron, it has more protons than electrons, resulting in a positively charged ion (cation).

13. **Which scientist discovered the electron?**

- (a) Ernest Rutherford (b) Albert Einstein
 (c) J.J. Thomson (d) Niels Bohr

Reason: J.J. Thomson discovered the electron in 1897 through his experiments with cathode rays.

14. **What is the purpose of the atomic number in the periodic table?**

- (a) To indicate the mass of the atom
 (b) To show the number of protons in an atom
 (c) To determine the chemical properties of an atom
 (d) To identify the isotope

Reason: The atomic number indicates the number of protons in an atom's nucleus, which determines the element's identity and its position in the periodic table.

15. **Which of the following statements is true about isotopes of the same element?**

- (a) They have different atomic numbers
 (b) They have the same number of neutrons
 (c) They have the same number of electrons
 (d) They have different chemical properties

Reason: Isotopes of the same element have the same number of protons (and electrons in a neutral atom) but differ in the number of neutrons.

ANSWERS

1.	(b)	2.	(c)	3.	(c)	4.	(c)	5.	(c)
6.	(c)	7.	(a)	8.	(c)	9.	(c)	10.	(b)
11.	(a)	12.	(b)	13.	(c)	14.	(b)	15.	(c)

Questions for Short Answers

i. **Why is it said that almost all the mass of an atom is concentrated in its nucleus?**

Ans. Almost all the mass of an atom is concentrated in its nucleus because the nucleus contains protons and neutrons, which have much greater mass than electrons. The mass of the electrons is negligible compared to that of protons and neutrons, so the majority of the atom's mass is in the nucleus.

ii. **Why are elements different from one another?**

Ans. Elements are different from one another because they have different numbers of protons in their atomic nuclei. The number of protons determines the element's atomic number, which defines the element and its unique chemical properties.

iii. **How many neutrons are present in ${}_{83}^{210}\text{Bi}$?**

Ans. The bismuth isotope ${}_{83}^{210}\text{Bi}$ has a mass number of 210 and an atomic number of 83 (indicating 83 protons). To find the number of neutrons, subtract the atomic number from the mass number:

$$210 - 83 = 127.$$

So, there are 127 neutrons in ${}_{83}^{210}\text{Bi}$.

iv. **Why is tritium (${}^3_1\text{H}$) a radioactive element?**

Ans. Tritium (${}^3_1\text{H}$) is a radioactive element because it has an unstable nucleus. It contains 2 neutrons and 1 proton, which makes its nucleus unstable. Tritium decays into helium -3 by emitting beta radiation, making it a radioactive isotope.

v. **How can an atom absorb and emit energy?**

Ans. An atom can absorb energy when electrons in lower energy levels move to higher energy levels (excited state) by absorbing energy from external sources, like heat or light. The atom emits energy when the electrons return to their lower energy levels (ground state), releasing the absorbed energy in the form of light or other types of radiation.

Additional Short Question and Answers

1. **What is the charge of an electron?**

Ans. The charge of an electron is negative.

2. **What are the three subatomic particles that make up an atom?**

Ans. The three subatomic particles are protons, neutrons, and electrons.

3. **Who discovered the electron?**

Ans. The electron was discovered by J.J. Thomson in 1897.

4. **What is an isotope?**

Ans. An isotope is an atom of the same element that has the same number of protons but a different number of neutrons.

5. **What is the atomic number of an element?**

Ans. The atomic number of an element is the number of protons in the nucleus of an atom.

6. **What does the mass number of an element represent?**

Ans. The mass number of an element is the sum of the number of protons and neutrons in the nucleus.

7. **Why do isotopes of an element have the same chemical properties?**

Ans. Isotopes have the same number of electrons and the same electron configuration, which is responsible for their chemical properties.

8. **What is the relative atomic mass of an element?**

Ans. The relative atomic mass of an element is the weighted average mass of all the isotopes of that element, relative to the mass of carbon-12.

9. **How do electrons exist in an atom according to Bohr's model?**

Ans. According to Bohr's model, electrons revolve around the nucleus in fixed, circular orbits or shells.

10. **What is the role of neutrons in the nucleus of an atom?**

Ans. Neutrons help stabilize the nucleus by offsetting the repulsive forces between positively charged protons.

11. **What happens when an atom gains or loses an electron?**

Ans. When an atom gains an electron, it becomes a negatively charged ion (anion), and when it loses an electron, it becomes a positively charged ion (cation).

12. **Why is hydrogen considered a unique element?**

Ans. Hydrogen is unique because it is the lightest element, having only one proton and one electron in its most common isotope.

13. **What is the significance of the atomic mass unit (amu)?**

Ans. The atomic mass unit (amu) is used to express the mass of atoms and subatomic particles. It is defined as one-twelfth the mass of a carbon-12 atom.

14. **What is the process of ionization?**

Ans. Ionization is the process in which an atom or molecule gains or loses electrons, resulting in the formation of ions.

15. **Why are radioactive isotopes useful in medicine?**

Ans. Radioactive isotopes are useful in medicine for diagnostic imaging, treatment of certain diseases, and radiotherapy, as they emit radiation that can be detected or used for treatment.

Constructed Response Questions

i. **Why does the energy of an electron increase as we move from the first shell to the second shell?**

Ans. The energy of an electron increases as we move from the first shell to the second shell because the second shell is farther from the nucleus. Electrons in higher energy levels are less tightly bound to the nucleus, requiring more energy to maintain their position.

ii. **Why is it needed to lower the pressure of the gas inside the discharge tube?**

Ans. Lowering the pressure of the gas inside the discharge tube reduces the number of collisions between gas particles, allowing the electrons to move more freely and easily ionize atoms, which makes it easier to observe the behavior of electrons.

iii. **What is the classical concept of an electron? How has this concept changed with time?**

Ans. The classical concept of an electron was that it is a negatively charged particle revolving around the nucleus in circular orbits. Over time, this concept has evolved into the quantum mechanical model, where electrons are described as existing in regions of probability around the nucleus, not fixed orbits.

iv. **Why are the nuclei of the radioactive elements unstable?**

Ans. The nuclei of radioactive elements are unstable because they contain an excess of energy or mass, causing the nucleus to be too large or have an imbalance between protons and neutrons. This instability leads to the emission of radiation as the nucleus tries to reach a more stable state.

v. **During discharge tube experiments, how did the scientists conclude that the same type of electrons and protons are present in all the elements?**

Ans. Scientists concluded that the same type of electrons and protons are present in all elements because, in discharge tube experiments, the same behavior (like deflection in electric and magnetic fields) was observed for the particles regardless of the gas inside the tube, indicating they were the same type of particles.

Descriptive Questions

i. **Explain the structure of a hydrogen atom.**

Ans. See Q. No.9

ii. **How does the theory of atomic structure explain the ionization of atoms by a radioactive isotope?**

Ans. See Q. No.10

iii. What is radioactivity? Explain any three applications of radioactive isotopes.

Ans. See Q. No.10,11

iv. Find out the relative atomic mass of mercury from the following data.

Isotope	Relative Abundance	Isotope	Relative Abundance
$^{196}\text{Hg} =$	0.0146%	^{199}Hg	= 16.34%
$^{198}\text{Hg} =$	10.02%	^{200}Hg	= 23.13%
$^{201}\text{Hg} =$	13.22%	^{202}Hg	= 29.80%
$^{204}\text{Hg} =$	6.85%		

Ans. To calculate the relative atomic mass of mercury (Hg) using the given data, we will use the formula for the weighted average:

$$\text{Relative Atomic Mass} = \sum (\text{Isotopic Mass} \times \text{Isotopic} \frac{\text{Abundance}}{100})$$

Given data:

Isotope	Relative Abundance	Isotope	Relative Abundance
$^{196}\text{Hg} =$	0.0146%	^{199}Hg	= 16.34%
$^{198}\text{Hg} =$	10.02%	^{200}Hg	= 23.13%
$^{201}\text{Hg} =$	13.22%	^{202}Hg	= 29.80%
$^{204}\text{Hg} =$	6.85%		

Now, we will multiply the isotopic masses by their respective abundances (in decimal form), and sum them up:

$$\begin{aligned} \text{Relative Atomic Mass} &= (196 \times 0.000146) + (198 \times 0.1002) + (199 \times 0.1634) \\ &+ (200 \times 0.2313) + (201 \times 0.1322) + (202 \times 0.2980) \\ &+ (204 \times 0.0685) \end{aligned}$$

$$\begin{aligned} \text{Relative Atomic Mass} &= (28.616) + (19.768) + (32.535) + (46.260) \\ &+ (26.595) + (60.202) + (13.978) \end{aligned}$$

$$\text{Relative Atomic Mass} \approx 227.954$$

Therefore, the relative atomic mass of mercury is approximately **227.95** amu.

Investigative Questions

i. How can scientists synthesize elements in the laboratory?

Ans. Scientists can synthesize elements in the laboratory by bombarding lighter elements with high-energy particles, such as protons or neutrons, to create heavier elements. This process often takes place in particle accelerators, where the atoms are forced to collide with each other, resulting in the formation of new, heavier elements.

ii. A system just like our solar system exists in an atom. Comment on this statement.

Ans. This statement refers to Bohr's model of the atom, where electrons revolve around the nucleus in fixed orbits, similar to how planets revolve around the sun. However, this analogy is limited because, in the atom, electrons are governed by quantum mechanics and do not follow exact circular paths as planets do, but instead exist in probability clouds around the nucleus.

Q.1. What is the reason behind the formation of chemical bonds in atoms?

Ans. Tendency to Lower Energy:

Atoms form chemical bonds to lower their energy, which makes them more stable.

Stability of Noble Gases:

The noble gases (He, Ne, Ar, Kr, Xe) have stable electron configurations, which explains why they do not react easily. Helium has two electrons, while other noble gases have eight electrons in their outermost shell, making them stable and unreactive.

Duplet and Octet Rule:

The Duplet Rule (for hydrogen and helium) and the Octet Rule (for other elements), state that atoms tend to be stable when they have two or eight electrons in their outermost shell.

Formation of Bonds for Stability:

Atoms form bonds by either gaining or losing electrons to complete their duplet or octet.

Example: sodium loses one electron to achieve a stable configuration, and hydrogen can either lose or gain one electron to achieve a stable duplet.

Electropositivity and Electronegativity:

Alkali and alkaline earth metals tend to lose electrons (electropositive), forming bonds with electronegative elements like oxygen and chlorine. However, further research has shown that the octet rule is not always the most important factor in bond formation.

Q.2. What is a chemical bond and what are the different types of chemical bonds?

Ans. Definition of Chemical Bond:

A chemical bond is a force of attraction between atoms that holds them together in the form of a molecule or a compound.

Formation of Chemical Bonds:

When atoms of different substances come close to each other, they can either attract or repel each other. If the attractive forces dominate the repulsion forces, the energy of the system decreases, causing the atoms to form a new molecule. If the repulsive forces are stronger, the atoms move apart without forming a bond.

Types of Bonds:

There are three main types of chemical bonds:

1. Ionic Bond:

This occurs when one atom donates electrons to another, resulting in the formation of positively and negatively charged ions that are attracted to each other.

2. Covalent Bond:

This bond forms when two atoms share electrons to achieve a stable electron configuration.

3. Coordinate Covalent Bond:

In this bond, both electrons in the shared pair come from the same atom, forming a bond with another atom.

Important Information!

The arrangement of electrons around the nucleus of an atom in shells and sub-shells is called electronic configuration.

Q.3. How is an ionic bond formed, and what are some examples of ionic compounds?

Ans. Formation of Ionic Bond:

An ionic bond is formed when atoms lose or gain electrons in order to acquire a more stable electronic configuration, similar to the nearest noble gas. This is driven by the tendency of atoms to achieve a stable outer electron shell.

Example: Sodium Chloride (NaCl):

When sodium (Na) reacts with chlorine (Cl), sodium loses one electron from its outermost shell and chlorine gains that electron. This results in the formation of a sodium ion (Na^+) and a chloride ion (Cl^-), each achieving the electron configuration of the nearest noble gas (Ne for sodium and Ar for chlorine).

	1st shell	2nd shell	3rd shell
$_{11}\text{Na}$	2	8	1
$_{17}\text{Cl}$	2	8	7

Fig (3.1): Electronic Configurations of Sodium and Chlorine

The positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-) are held together by electrostatic forces, forming an ionic bond and creating sodium chloride (NaCl), an ionic compound.

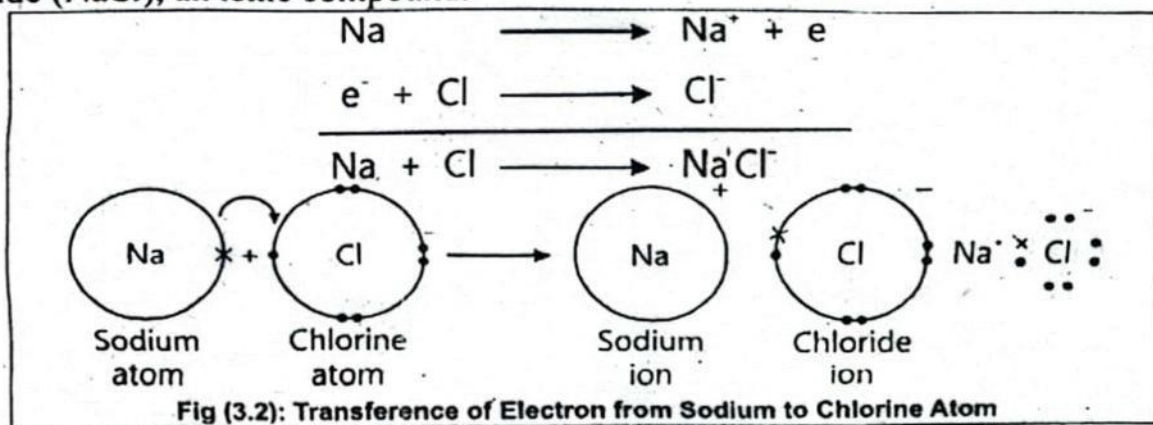


Fig (3.2): Transference of Electron from Sodium to Chlorine Atom

Other Examples:

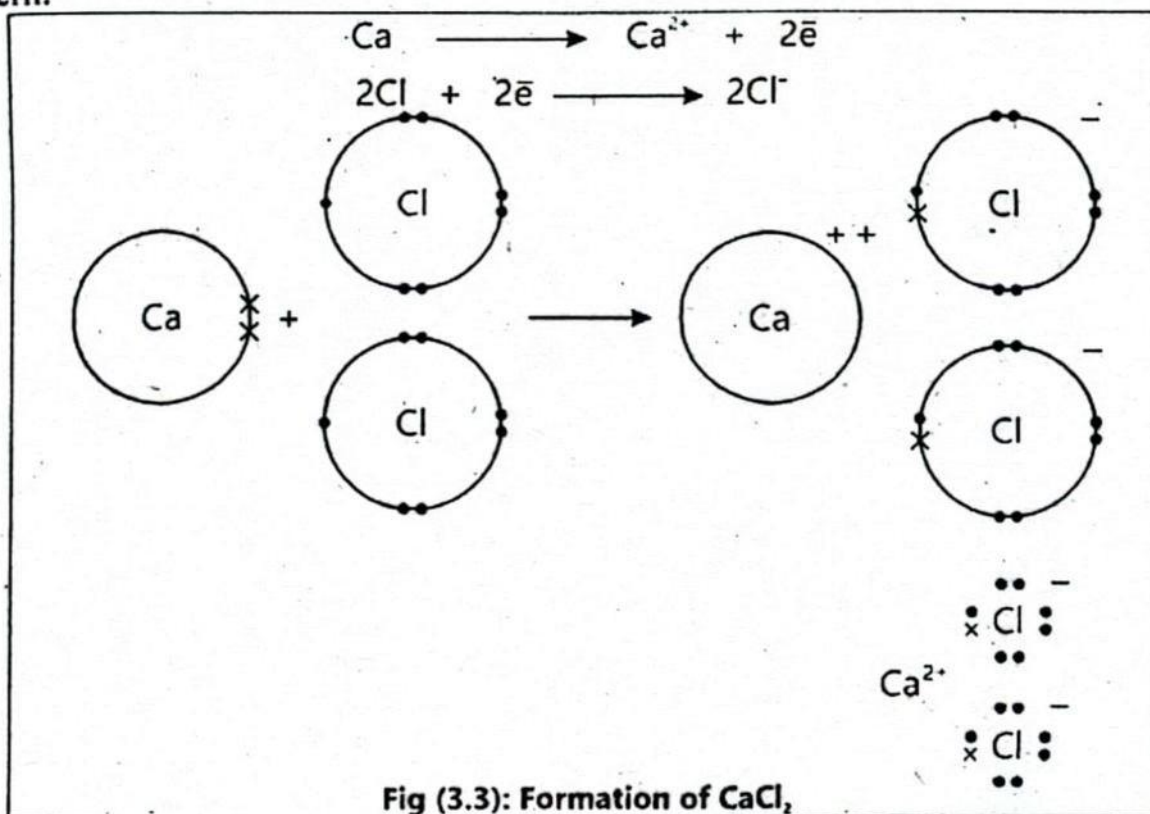
Sodium also reacts with fluorine to form sodium fluoride (NaF) and with bromine to form sodium bromide (NaBr).

Electrons Involved:

The electrons that take part in the formation of the ionic bond come only from the outermost electron shell of the atoms involved in the reaction.

Example: Calcium Chloride (CaCl_2):

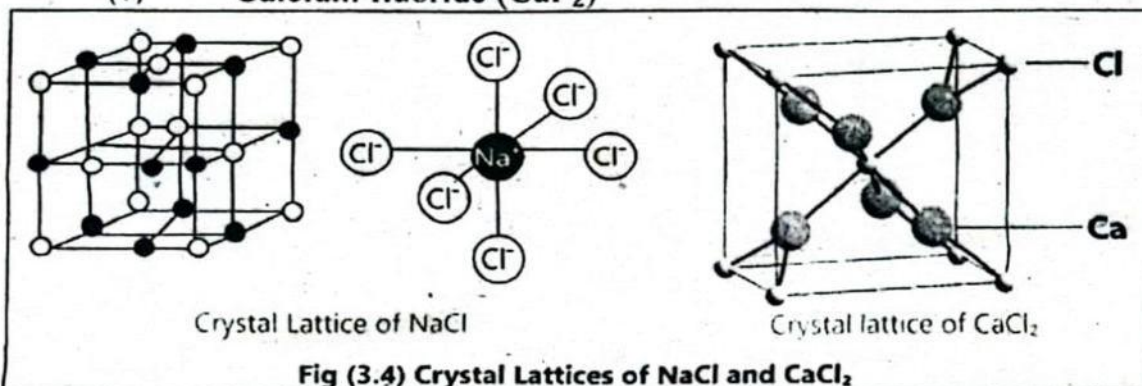
Calcium (Ca) loses two electrons to form a calcium ion (Ca^{2+}), and these ions combine with chloride ions (Cl^-) to form calcium chloride (CaCl_2). This results in the formation of a crystal lattice structure, where ions are arranged in a three-dimensional pattern.



Ionic Compounds:

Examples of ionic compounds include:

- (i) Potassium chloride (KCl)
- (ii) Magnesium fluoride (MgF_2)
- (iii) Sodium fluoride (NaF)
- (iv) Potassium bromide (KBr)
- (v) Calcium fluoride (CaF_2)



Figures of crystal lattices of NaCl and CaCl_2 . (Fig 3.4)

An ionic bond is therefore a bond which is formed by the complete transference of electron or electrons from one atom to another atom.

EXERCISE

1. What types of elements form ionic bonds?

Ans. Ionic bonds are typically formed between metals and non-metals. Metals (such as sodium, calcium) lose electrons to form positive ions, while non-metals (such as chlorine, oxygen) gain electrons to form negative ions.

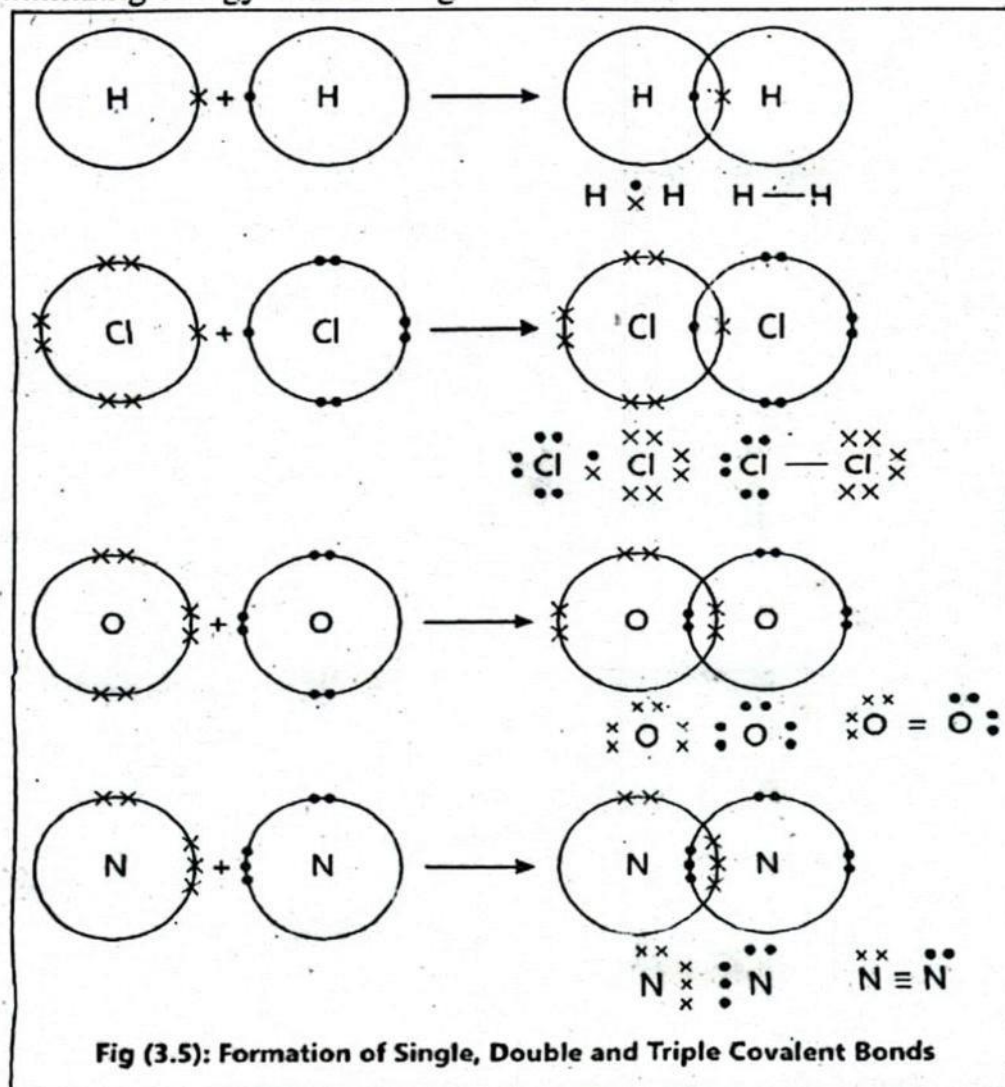
2. What are the conditions for an ionic bond to be formed?

Ans. An ionic bond forms when one atom (usually a metal) loses electrons, and another atom (usually a non-metal) gains those electrons. This results in the formation of oppositely charged ions that are held together by electrostatic attraction.

Q.4. How is a covalent bond formed, and what are the different types of covalent bonds?

Ans. **Formation of Covalent Bond:**

A covalent bond forms when two atoms share electrons to achieve a more stable electron configuration. As the atoms approach each other, their electrons come under the influence of the nucleus of the other atom, creating an attractive force. At the same time, there is repulsion between the electrons and nuclei of the atoms. The atoms will settle at a distance where the attractive forces dominate the repulsive forces, minimizing energy and forming a stable molecule.



Single Covalent Bond:

A single covalent bond is formed when two atoms share one pair of electrons. It is represented by a single line (-) between the atoms.

Double Covalent Bond:

In a double covalent bond, two atoms share two pairs of electrons. This is represented by two lines (=) between the atoms.

Triple Covalent Bond:

A triple covalent bond occurs when two atoms share three pairs of electrons. It is represented by three lines (\equiv) between the atoms.

Representation of Electrons:

The shared electrons can be shown using dots or crosses in the diagram of the molecule. Examples of covalent bonding are depicted in figures showing single, double, and triple bonds.

EXERCISE

1. What type of elements form covalent bonds?

Ans. Covalent bonds are typically formed between non-metals. These atoms share electrons to achieve a stable electron configuration.

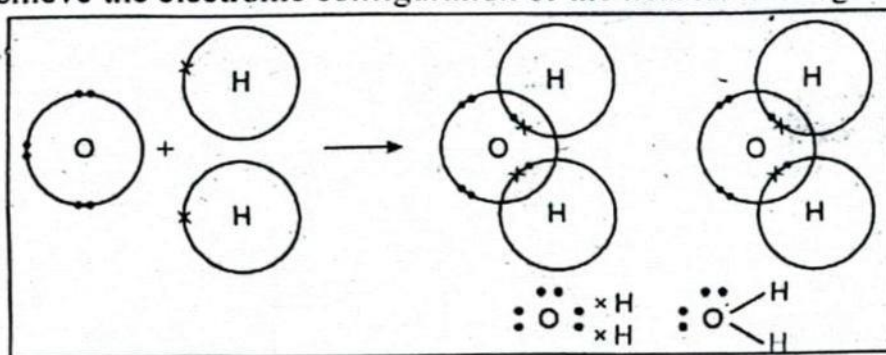
2. How is a covalent bond different from an ionic bond?

Ans. In a covalent bond, electrons are shared between atoms, while in an ionic bond, electrons are transferred from one atom to another, resulting in the formation of oppositely charged ions.

Q.5. How are covalent compounds like water and carbon dioxide formed?

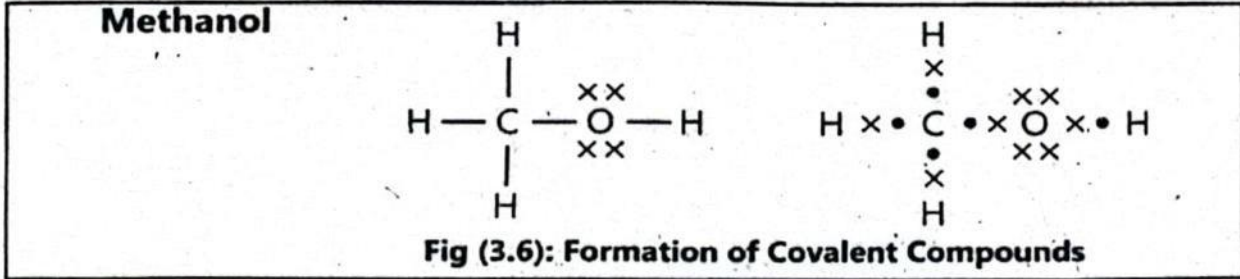
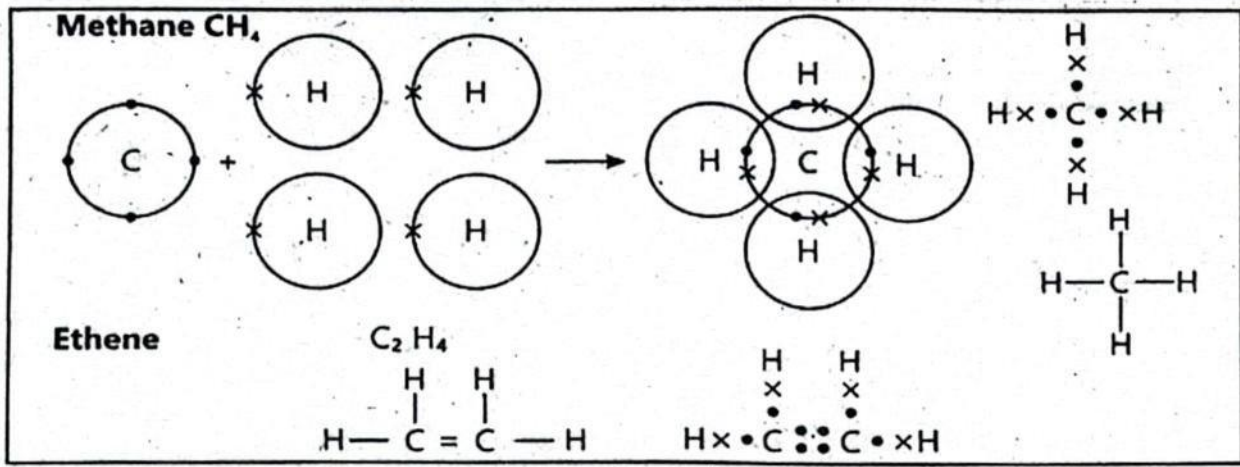
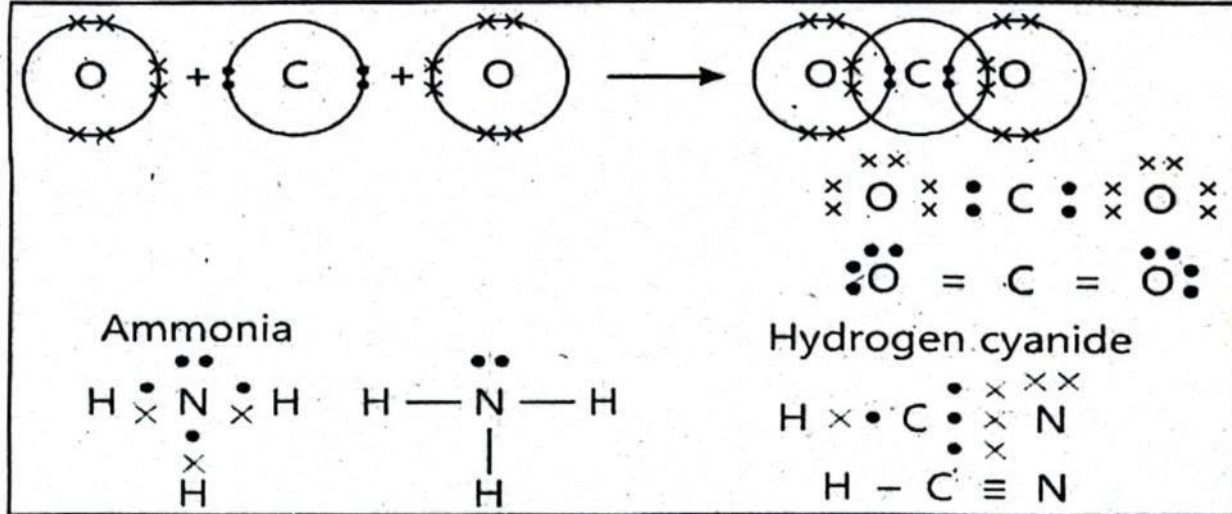
Ans. Formation of Water:

A water molecule (H_2O) is formed when two hydrogen atoms each share one electron with an oxygen atom. This results in the sharing of electrons between the hydrogen and oxygen atoms, forming a stable molecule. After sharing, both hydrogen and oxygen achieve the electronic configuration of the nearest noble gases.



Formation of Carbon Dioxide:

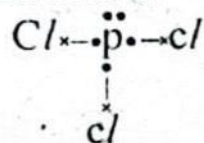
A carbon dioxide molecule (CO_2) is formed when a carbon atom shares its four electrons with two oxygen atoms. Each oxygen atom shares two electrons with the carbon atom, resulting in a stable molecule. After electron sharing, both carbon and oxygen atoms attain the electronic configuration of the nearest noble gases.



EXERCISE

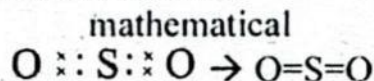
- i. Draw electron dot and cross structure of the following compounds.
SiH₄, PCl₃, SO₂, SO₃
- Ans. Here are the electron dot and cross structures for the given compounds:
1. **SiH₄ (Silane):**
Silicon (Si) shares four electrons with four hydrogen (H) atoms.
Electron dot and cross structure:
- $$\begin{array}{c} \text{H} \\ \times \\ \cdot \\ \text{H} \times \cdot \text{Si} \cdot \times \text{H} \\ \cdot \\ \times \\ \text{H} \end{array}$$

2. **PCl₃ (Phosphorus trichloride):**
Phosphorus (P) shares three electrons with three chlorine (Cl) atoms.
Electron dot and cross structure:

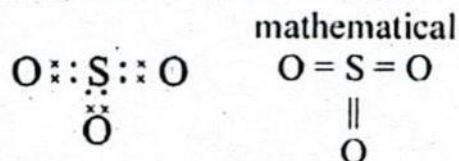


(Lone pair on P)

3. **SO₂ (Sulfur dioxide):**
Sulfur (S) shares two electrons with two oxygen (O) atoms.
Electron dot and cross structure:



4. **SO₃ (Sulfur trioxide):**
Sulfur (S) shares three electrons with three oxygen (O) atoms.
Electron dot and cross structure:



In these structures, the dots and crosses represent the shared electrons between atoms.

Q.6. What is the difference between polar and non-polar covalent bonds?

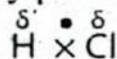
Ans. Non-Polar Covalent Bond:

A non-polar covalent bond is formed when two identical atoms share electrons equally. The shared pair of electrons lies in the middle of the two atoms, resulting in no partial charges. An example is the covalent bond in molecules like hydrogen (H₂) and chlorine (Cl₂), which are non-polar.

Polar Covalent Bond:

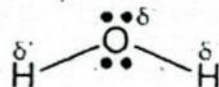
Explanation:

A polar covalent bond occurs when two non-identical atoms form a bond and the shared electrons are attracted more towards the atom with higher electronegativity. This creates a partial negative charge on the more electronegative atom and a partial positive charge on the less electronegative atom. An example is the hydrogen chloride (HCl) molecule, where the shared electrons are more attracted to chlorine, making it partially negative and hydrogen partially positive.



Examples of Polar Covalent Bonds:

In water (H₂O), the electrons are shared unequally between hydrogen and oxygen, resulting in a polar covalent bond.



Q.7. What is a coordinate covalent bond, and how is it formed?

Ans. Definition of Coordinate Covalent Bond:

A coordinate covalent bond is a type of covalent bond where both electrons in the shared pair are donated by only one atom.

Explanation:

This bond forms when an atom with a lone pair of electrons donates them to another atom or species that does not have a full electron pair to share.

Donor and Acceptor:

The atom that donates the electron pair is called the **donor**, and the atom that accepts the electron pair is called the **acceptor**.

Representation:

The coordinate covalent bond is represented by an arrow (\rightarrow) pointing towards the acceptor atom.

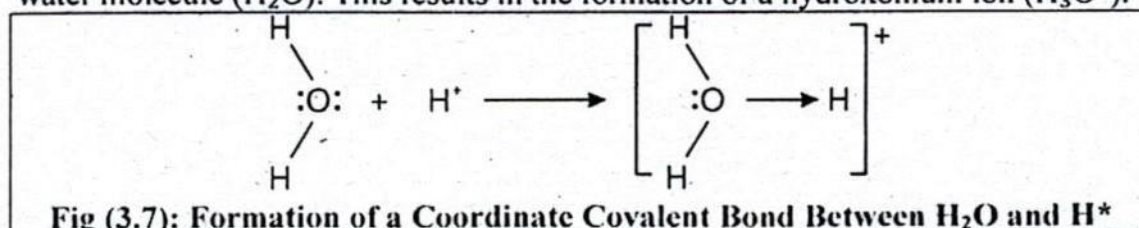
Example:

One example of a coordinate covalent bond is the formation of ammonium ion (NH_4^+), where the nitrogen atom (with a lone pair of electrons) donates its electron pair to a proton (H^+), forming the bond.

Q.8. How is a hydroxonium ion (H_3O^+) formed, and what is its nature?

Ans. Formation of Hydroxonium Ion (H_3O^+):

When an acid dissolves in water, it provides protons (H^+). The proton (H^+) has an empty outer shell and can accept one of the electron pairs from the oxygen atom in a water molecule (H_2O). This results in the formation of a hydroxonium ion (H_3O^+).



Coordinate Covalent Bond:

The bond formed between the water molecule and the proton is a **coordinate covalent bond**, where the oxygen atom in water donates both electrons in the bond to the proton.

Nature of the Hydroxonium Ion:

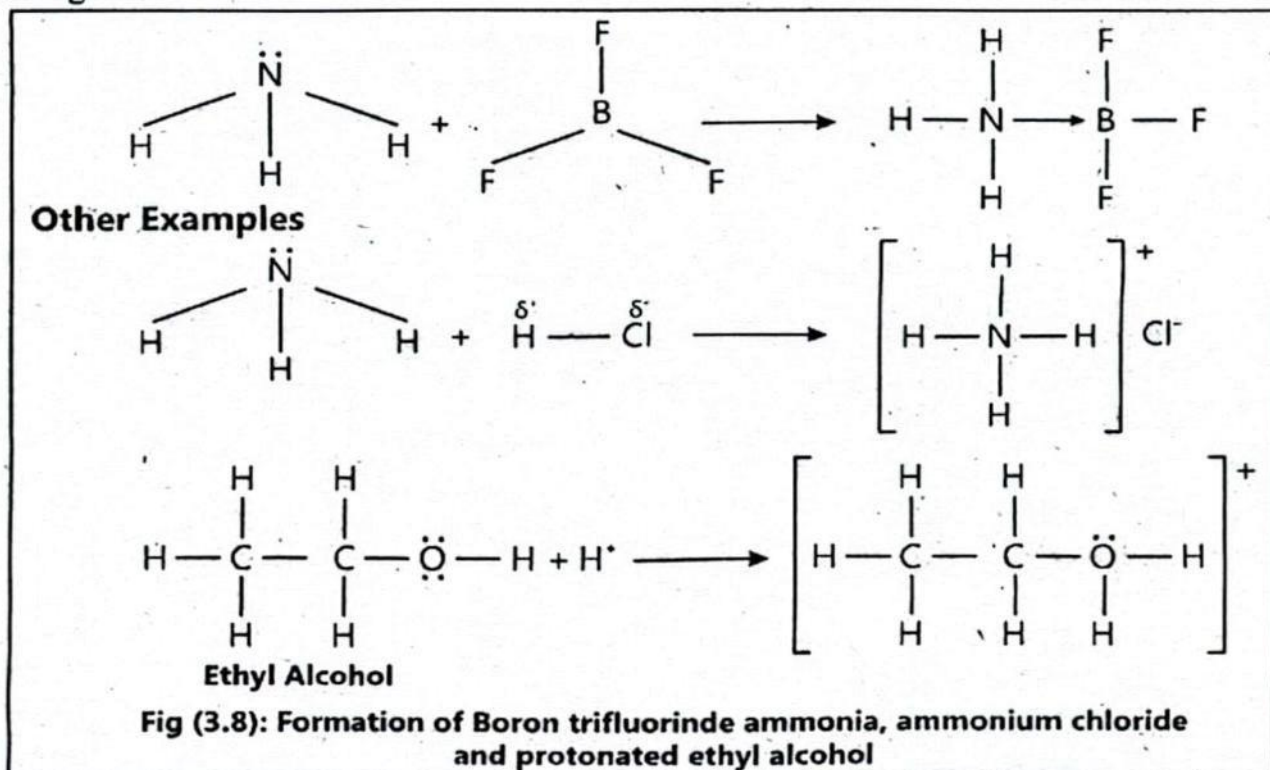
Once the hydroxonium ion is formed, the three bonds between the oxygen and hydrogen atoms behave similarly, and there is no significant difference between a coordinate covalent bond and a regular covalent bond. The positive charge is distributed over the entire hydroxonium ion.

Q.9. How is a coordinate covalent bond formed between ammonia and boron tri fluoride?

Ans. Reaction Between Ammonia (NH_3) and Boron Trifluoride (BF_3):

In the reaction between ammonia (NH_3) and boron trifluoride (BF_3), ammonia acts as a donor of an electron pair. The nitrogen atom in ammonia, which has a lone pair of

electrons, donates one pair to fill the partially empty outer shell of boron in boron trifluoride. This results in the formation of a **coordinate covalent bond** between nitrogen and boron.



Nature of the Bond:

The bond formed is a coordinate covalent bond, where both electrons come from the nitrogen atom in ammonia, and the boron atom in BF_3 accepts the electron pair. After the bond is formed, there is no difference between a regular covalent bond and a coordinate covalent bond; both behave similarly.

Bond Breaking:

When a coordinate covalent bond breaks, both of the bonded electrons are transferred back to the donor species (in this case, ammonia).

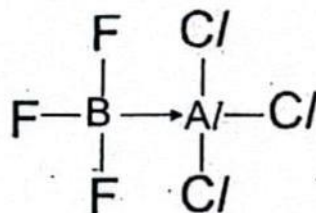
EXERCISE

1. Draw the pictures of coordinate covalent bond formed between:

(a) BF_3 and AlCl_3

In this case, boron trifluoride (BF_3) donates an electron pair to aluminum chloride (AlCl_3), forming a coordinate covalent bond.

• **Electron Dot Structure:**



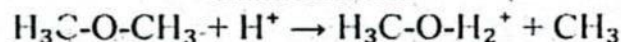
(b) CH_3OCH_3 and H^+

In this case, the oxygen in dimethyl ether (CH_3OCH_3) donates an electron

pair to a proton (H^+), forming a coordinate covalent bond.

• **Electron Dot Structure:**

mathematical



2. **Which compound is not able to form a coordinate covalent bond?**

Ans. **Methane (CH_4)** is not able to form a coordinate covalent bond because it does not have a lone pair of electrons to donate.

Q.10. **Explain the nature of metallic bonding and how it accounts for the properties of metals.**

Ans. **Nature of Metallic Bonding:**

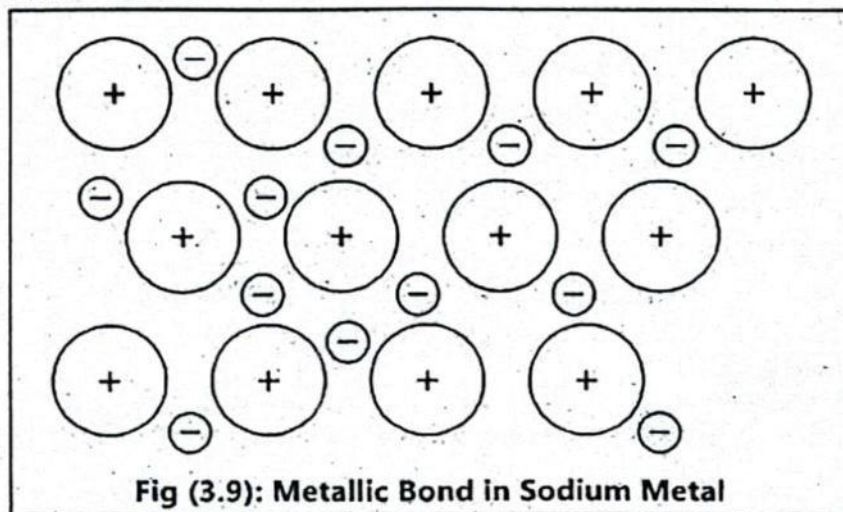
Metallic bonding is the type of chemical bond that holds metals together. In a metal, the atoms are arranged in a closely packed structure, and their outermost electrons are free to move throughout the entire metal structure. This leads to the formation of what is known as a "sea of electrons" that is shared by all the positively charged metal ions.

Properties of Metals Explained by Metallic Bonding:

1. **Metallic Lustre:**

Metals are known for their shiny appearance, or metallic lustre. This property arises because the free electrons in the metal surface can reflect light efficiently, giving metals their characteristic shine.

Fig 3.9 shows the structure of sodium metal where the mobile electrons reflect light and give metallic lustre.

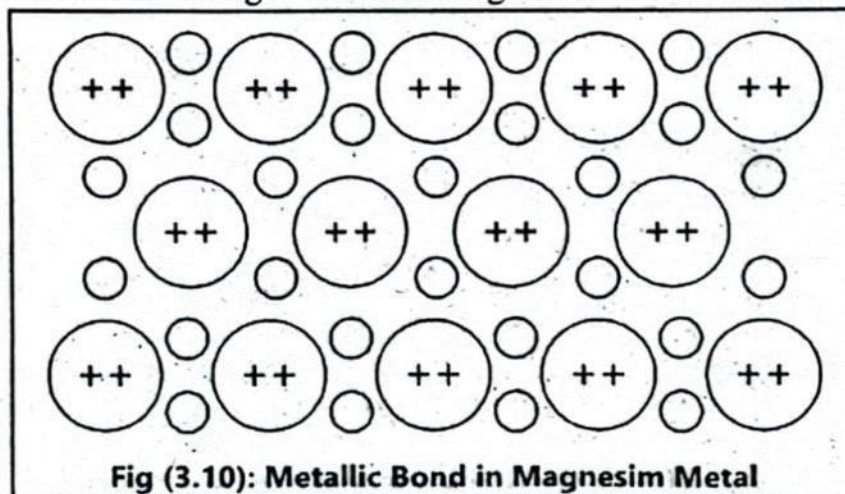


2. **High Melting and Boiling Points:**

Metals typically have high melting and boiling points because the metallic bonds are strong, requiring a lot of energy to break the bond between the metal ions and the electron sea.

The strength of the metallic bond depends on how many electrons each atom contributes to the electron sea and how strongly the ions are attracted to the electrons. For example, **sodium metal (Na)** has weaker bonds because it only releases one electron, while **magnesium metal (Mg)** releases two electrons, creating a stronger bond and higher melting/boiling points.

Fig 3.10 illustrates the difference between sodium and magnesium, showing that magnesium's bond is stronger due to the higher number of mobile electrons.



3. Good Conductors of Heat and Electricity:

Since the electrons in metals are free to move, they can conduct electricity by moving in response to an electric field. Similarly, when heat is applied, the free electrons absorb and transfer the thermal energy quickly throughout the metal.

This is why metals are excellent conductors of both electricity and heat.

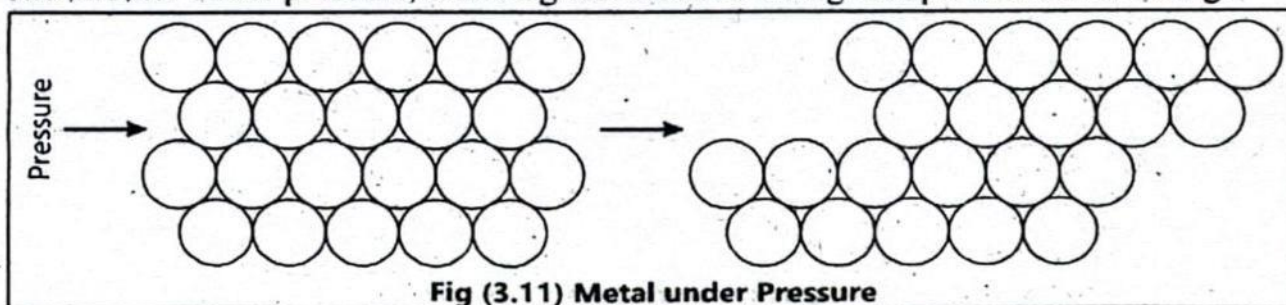
4. Hard and Heavy:

Metals are usually dense and hard due to the strong metallic bonds that hold the atoms tightly together. The closely packed arrangement of metal ions in a regular pattern also contributes to the hardness and high density of metals.

5. Malleability and Ductility:

Metals can be easily shaped and stretched into wires or sheets because the layers of atoms can slide over each other without breaking the metallic bond. This ability is a result of the free movement of electrons that helps maintain the bond even when the metal is deformed.

Fig 3.11 illustrates the arrangement of atoms in a metal and how they can slip past one another under pressure, allowing the metal to change shape without breaking.



Conclusion:

Metallic bonding explains the distinct physical properties of metals. The "sea of electrons" allows metals to conduct heat and electricity, provides metallic lustre, and contributes to their hardness and strength. The strength of metallic bonds varies based on the number of mobile electrons and the charge of the metal ions, influencing the metal's melting point, hardness, and ductility.

Interesting Information!

Metals are extensively used in many industries. They are used in machinery, automobiles, railways, air crafts, rockets, in construction industry, in electronics industry, in jewelry, in electric wires and many more.

EXERCISE

1. **What type of atoms form metallic bond?**

Ans. Metallic bonds are formed between metal atoms. These atoms have low ionization energies, allowing their outer electrons to be easily lost and shared in a "sea of electrons" that move freely throughout the metal.

2. **Give a comparison of metallic bond with an ionic bond.**

Ans. Metallic Bond: Metallic Bond Involves the attraction between positively charged metal ions and a sea of delocalized electrons. The bond is non-directional and allows metal atoms to conduct electricity, be malleable, and have high melting points.

Ionic Bond: Ionic Bond Formed by the transfer of electrons from one atom (usually a metal) to another (usually a non-metal), creating positively and negatively charged ions. The bond is strong but directional, and ionic compounds typically have high melting points and conduct electricity when molten or dissolved in water.

Q.11. **What is the electropositive character of metals, and how does it vary among different types of metals?**

Ans. Electropositive Character of Metals:

Metals have a tendency to lose electrons and form positive ions (cations). This tendency is known as the **electropositive character** of metals. The more easily a metal can lose electrons, the more electropositive it is.

Variation Among Different Metals:

1. **Alkali Metals (Na, K):**

Alkali metals are highly electropositive and lose their outer electrons easily. As a result, they are very reactive. They react vigorously with water, halogens, and acids to form hydroxides, halides, salts, and hydrogen.

2. **Alkaline Earth Metals (Mg, Ca):**

Alkaline earth metals are less electropositive than alkali metals because they lose their outer electrons less easily. Their reactions with water and halogens are also less vigorous.

3. **Aluminum:**

Aluminum is also highly electropositive and reacts readily with mineral acids to form salts and hydrogen.

Q.12. **What is the electronegative character of non-metals, and how do they behave in chemical reactions?**

Ans. Electronegative Character of Non-metals:

Non-metals have a strong affinity for electrons. They tend to gain electrons easily and form negatively charged ions called **anions**. This tendency makes them **electronegative elements**.

Fluorine is the most electronegative element in the periodic table, followed by **oxygen, nitrogen, and chlorine.**

Behavior in Chemical Reactions:

Non-metals readily react with metals to form **ionic bonds**, where they gain electrons from metals, resulting in the formation of anions.

Non-metals also combine with other non-metals to form a wide variety of **molecular substances**, typically through **covalent bonding.**

Q.13. Compare the properties of ionic and covalent compounds.

Property	Ionic Compounds	Covalent Compounds
1. Structure	Ionic Compounds are Composed of oppositely charged ions arranged in a crystalline structure. The compound is neutral, with a strong electrostatic force between ions.	Covalent Compounds Exist as discrete neutral molecules with strong electrostatic attraction between the nuclei and shared electrons.
2. Melting and Boiling Points	High melting and boiling points due to strong electrostatic forces between ions. Example: Sodium chloride melts at 801°C.	Generally lower melting and boiling points compared to ionic compounds. Low molecular mass compounds may be gases or liquids, while high molecular mass compounds may be solids.
3. Solubility	Generally soluble in polar solvents like water.	Usually insoluble in water but soluble in non-polar solvents like ether, benzene, and acetone.
4. Electrical Conductivity	Good conductors of electricity in molten state or in solution form due to free ions.	Poor conductors of electricity as they lack free ions.

This comparison highlights the key differences in properties between ionic and covalent compounds.

Q.14. What are intermolecular forces of attraction, and how do they influence the physical properties of substances?

Ans. Intermolecular Forces of Attraction

Intermolecular forces are the attractive forces that exist between the molecules of elements and compounds. These forces are generally much weaker than the bonding forces within the molecules themselves.

Effect on Physical Properties

Melting and Boiling Points:

The strength of the intermolecular forces directly influences the physical properties of a substance.

Stronger Forces: Higher melting and boiling points, as it requires more energy to overcome the stronger attractions between molecules.

Weaker Forces: Lower melting and boiling points, as less energy is needed to separate the molecules.

In summary, intermolecular forces significantly affect how substances behave in different states of matter and their thermal properties.

Q.15. What are dipole-dipole forces of attraction, and how do they influence the physical properties of compounds?

Ans. Dipole-Dipole Forces of Attraction

Dipole-dipole forces occur between the molecules of polar compounds. These forces arise because of the unequal sharing of electrons between atoms with different electronegativities, creating partial positive and partial negative charges.

Example: Example, in the HCl molecule, chlorine has a greater electronegativity than hydrogen, causing it to attract the shared electrons more strongly. This creates a polar bond, where chlorine carries a partial negative charge (δ^-) and hydrogen carries a partial positive charge (δ^+).



Effect on Physical Properties

Higher Melting and Boiling Points:

Substances with dipole-dipole forces typically have higher melting and boiling points compared to those without such forces. This is because the dipole-dipole attractions between molecules require more energy to overcome when transitioning from solid to liquid or liquid to gas.

Example: In HCl, the molecules attract each other due to these dipole-dipole forces, which influence its physical properties like its higher boiling point compared to non-polar molecules.

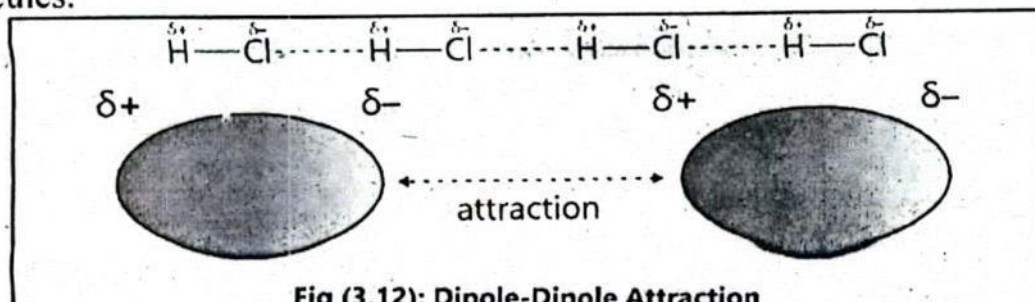


Fig (3.12): Dipole-Dipole Attraction

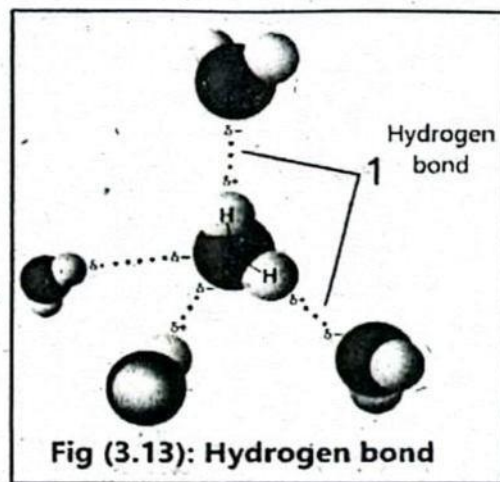
Figure 3.12 illustrates how dipole-dipole forces cause the molecules of HCl to attract each other, demonstrating the nature of these intermolecular forces.

Q.16. What is hydrogen bonding, and how does it affect the physical properties of compounds?

Ans. Hydrogen Bonding

Hydrogen bonding is a special type of dipole-dipole attraction that occurs when hydrogen is covalently bonded to highly electronegative elements such as fluorine (F), oxygen (O), or nitrogen (N). The significant difference in electronegativity between hydrogen and these elements makes the covalent bond highly polar, leading to strong dipole-dipole attractions between molecules.

For instance, in water (H_2O), the O—H bonds are highly polar due to the electronegativity difference between oxygen and hydrogen. This polarity creates strong attractive forces between water molecules, known as **hydrogen bonding**.



Effect on Physical Properties

Higher Melting and Boiling Points:

The strength of hydrogen bonds explains why water has relatively higher melting and boiling points compared to other compounds like H_2S and NH_3 , which do not exhibit as strong hydrogen bonding.

Example:

Water (H_2O) molecules exhibit hydrogen bonding, resulting in its high boiling and melting points compared to other molecules with similar molecular masses but weaker intermolecular forces.

Figure 3.13 illustrates how hydrogen bonding occurs between water molecules, showing the strength and effect of these attractive forces.

Q.17. What are the properties of ionic compounds and how do they behave in different states and solutions?

Ans. Properties of Ionic Compounds:

1. Crystal Lattice Structure:

Ionic compounds are composed of oppositely charged ions arranged in a crystal lattice. The electrostatic attraction between these ions holds the compound together in a stable, rigid structure. The bonds between the ions are non-directional, allowing the ions to surround each other in all directions.

2. High Melting and Boiling Points:

The strong electrostatic forces between the ions require significant energy to break, resulting in high melting and boiling points for ionic compounds.

3. Brittleness:

Ionic compounds are brittle. When an external force is applied to the crystal lattice, it causes the lattice to break easily because like charges repel each other, which causes the compound to fracture.

4. **Electrical Conductivity:**

In the solid state, ionic compounds do not conduct electricity because the ions are fixed in place within the crystal lattice and cannot move. However, when ionic compounds are melted or dissolved in water, the ions become free to move and conduct electricity. This is due to the **electrolytic conductivity** in the molten or aqueous state.

5. **Solubility in Water:**

Ionic compounds are generally soluble in water. Water molecules break the electrostatic forces of attraction between the ions, and the resulting free ions are hydrated by the water molecules. This process is shown in **Figure 3.14**, which illustrates how water molecules surround and interact with the free ions.

6. **Reactivity in Aqueous Solutions:**

Ionic compounds react in aqueous solutions. For example, when solutions of silver nitrate (AgNO_3) and sodium chloride (NaCl) are mixed, a white precipitate of silver chloride (AgCl) forms due to the reaction between the positive ions of one compound and the negative ions of the other.

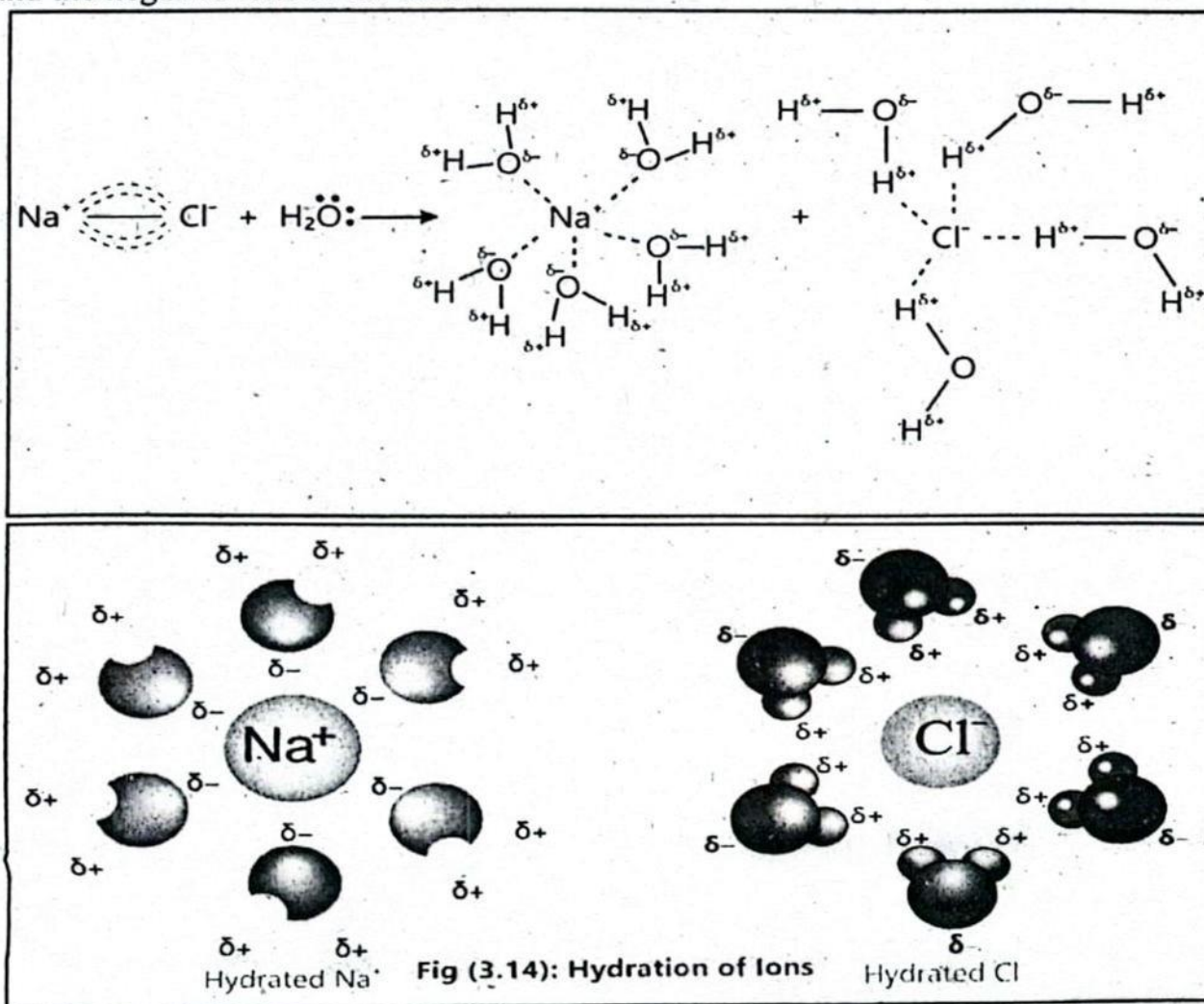
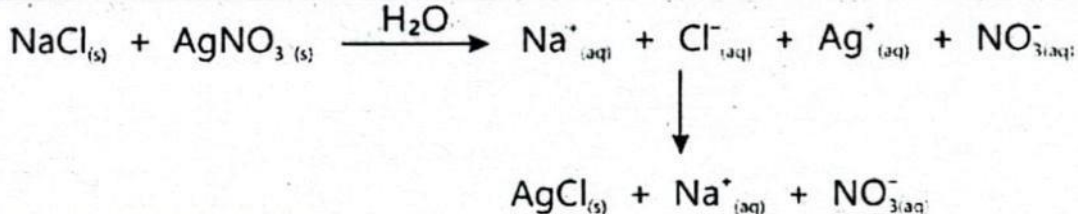


Figure 3.14 shows the process of hydration, where water molecules surround and interact with the free ions after the ionic compound dissolves in water.



Interesting Information!

Conduction of ionic compounds in molten state and in form of an aqueous solution has been utilized to prepare many important elements and compounds. For example, electrolysis of molten sodium chloride gives us sodium metal and chlorine gas. Similarly electrolysis of aqueous sodium chloride gives sodium hydroxide and chlorine gas.

Diamonds, due to their exceptional hardness, are highly valued in industries. Diamond tipped glass cutters are used to make clean cuts in glass. Diamond-tipped drill bits are used to drill through hard rocks in mining operation.

Graphite is used in pencils, in polishes and to make crucibles. Graphite electrodes are used in battery cells and in electric arc furnaces to produce steel.

Q.18. How do covalent compounds behave compared to ionic compounds? Discuss their physical properties, examples of solid forms, and behavior in water.

Ans.1. **Covalent Compounds vs. Ionic Compounds:**

Covalent Compounds: In covalent compounds, atoms are bonded by sharing electrons. These compounds behave differently from ionic compounds due to the weaker forces of attraction between molecules. They often have low boiling and melting points, and their physical properties can vary significantly depending on the structure of the compound.

Ionic Compounds: In ionic compounds, ions are held together by strong electrostatic forces, resulting in high melting points, brittleness, and the ability to conduct electricity in molten or aqueous states.

2. Physical Properties of Covalent Compounds:

Low Boiling and Melting Points: Most covalent compounds have low boiling and melting points due to the weak intermolecular forces. For example, diatomic molecules such as nitrogen (N_2), oxygen (O_2), fluorine (F_2), and chlorine (Cl_2) exist as gases at room temperature.

State of Matter:

Volatile Liquids: Some covalent compounds, like bromine (Br_2), exist as volatile liquids.

Covalent Solids: Elements such as carbon, phosphorus, and sulfur exist as covalent solids, which can be in either amorphous or crystalline forms. For example, coal is an amorphous form of carbon, while diamond and graphite are crystalline forms of carbon.

3. Examples of Covalent Solids:

Diamond: In diamond, each carbon atom is bonded to four other carbon atoms in a very strong and rigid structure, making diamond the hardest known substance. It is used for cutting, polishing, and drilling tools.

Graphite: Graphite has a layered structure of hexagonal carbon rings. The layers can slide past each other, making graphite an effective lubricant. The mobile electrons between the layers make graphite a good conductor of electricity, and it is used as an electrode.

4. Behavior of Binary Covalent Compounds:

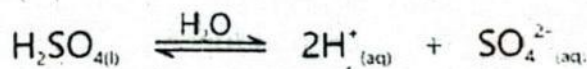
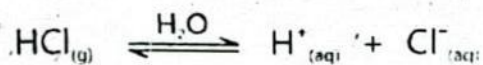
Gases at Room Temperature: Many binary covalent compounds exist as gases at room temperature, such as methane (CH_4), ammonia (NH_3), hydrogen chloride (HCl), carbon dioxide (CO_2), and sulfur dioxide (SO_2).

Liquids at Room Temperature: Water (H_2O) and hydrogen fluoride (HF) are exceptions and exist as liquids at room temperature. Water has a relatively high boiling point due to strong hydrogen bonds between molecules.

5. Covalent Compounds in Water:

Ionization in Water: Some covalent compounds, such as hydrogen chloride (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3), ionize completely in water, behaving as strong acids.

Intermolecular Forces in Water: Water has strong intermolecular forces (hydrogen bonding), which result in a high boiling point and make it a liquid at room temperature.



KEY POINTS

1. Why do atoms form bonds with other atoms?

Ans. Atoms form bonds with other atoms to stabilize themselves by obeying the **duplet rule** (for small atoms like hydrogen) and the **octet rule** (for larger atoms), achieving a stable electron configuration.

2. What is a chemical bond?

Ans. A chemical bond is the force of attraction that keeps atoms together in a compound or molecule.

3. What is an ionic bond?

Ans. An ionic bond is formed by the transfer of one or more electrons from one atom to another, resulting in positively and negatively charged ions.

4. How is a covalent bond formed?

Ans. A covalent bond is formed by the mutual sharing of electrons between atoms. It can be **single**, **double**, or **triple**, depending on the number of shared electron pairs.

5. What is a non-polar covalent bond?

Ans. A non-polar covalent bond is formed when a covalent bond occurs between two identical atoms with equal electronegativities, resulting in an equal sharing of electrons.

6. What is a polar covalent bond?

Ans. A polar covalent bond is formed when a covalent bond occurs between two non-identical atoms with different electronegativities, leading to an unequal sharing of electrons.

7. What are the types of intermolecular forces?

Ans. Intermolecular forces are of two main types:

- **Dipole-dipole attractions:** Present between molecules with permanent dipoles.
- **Hydrogen bonding:** A strong type of dipole-dipole interaction occurring when hydrogen is bonded to highly electronegative atoms like F, O, or N.

8. What is a coordinate covalent bond?

Ans. A coordinate covalent bond is formed when an electron pair is donated by one atom, ion, or molecule to form a bond with another atom or ion.

9. What are the properties of ionic solids?

Ans. Ionic solids are crystalline, have high melting and boiling points, are generally soluble in water, and conduct electricity in the molten state or aqueous solution.

10. What are the properties of covalent compounds?

Ans. Lower molecular mass covalent compounds are gases or low-boiling liquids.

(i) Higher molecular mass covalent compounds exist as solids.

(ii) They are poor conductors of electricity and are soluble in organic solvents like benzene or ether.

11. How are the properties of ionic and covalent compounds explained?

Ans. The properties of ionic and covalent compounds are explained by the type of attractive forces present:

- (i) **Ionic Compounds:** Strong electrostatic forces between ions lead to high melting points, solubility in water, and conductivity in molten or aqueous states.
- (ii) **Covalent Compounds:** Weaker intermolecular forces result in lower melting and boiling points, lack of conductivity, and solubility in organic solvents.

EXERCISE

1. Tick (✓) the correct answer.

i. When molten copper and molten zinc are mixed together, they give rise to a new substance called brass. Predict what type of bond is formed between copper and zinc.

- (a) Coordinate covalent bond (b) Ionic bond
(c) Metallic bond (d) Covalent bond

Reason: Brass is an alloy of copper and zinc where a metallic bond is formed between the metal atoms.

ii. Which element is capable of forming all the three types of bonds; covalent, coordinate covalent and ionic?

- (a) Carbon (b) Oxygen
(c) Magnesium (d) Silicon

Reason: Carbon can form covalent bonds (e.g., CH_4), coordinate covalent bonds (e.g., CO), and ionic bonds (e.g., CaC_2).

iii. Why is H_2O a liquid while H_2S is a gas?

- (a) Because in water, the atomic size of oxygen is smaller than that of sulphur
(b) Because water is a polar compound and there exists strong forces of attraction between its molecules
(c) Because H_2O molecule is lighter than H_2S
(d) Because water can easily freeze into ice

Reason: Water has strong hydrogen bonding, making it a liquid, while H_2S lacks such strong intermolecular forces and exists as a gas.

iv. Which of the following bonds is expected to be the weakest?

- (a) C-C (b) Cl-Cl
(c) O-O (d) F-F

Reason: Fluorine atoms are small, and the F-F bond is weakened due to strong repulsion between the lone pairs of electrons.

v. Which form of carbon is used as a lubricant?

- (a) Coal (b) Diamond
(c) Graphite (d) Charcoal

Reason: Graphite has a layered structure, and the weak forces between the layers make it suitable as a lubricant.

vi. Keeping in view the intermolecular forces of attraction, indicate which compound has the highest boiling point.

- (a) H_2O (b) H_2S
(c) HF (d) NH_3

Reason: Water has strong hydrogen bonding compared to H_2S , HF , and NH_3 , giving it the highest boiling point.

vii. Which metal has the lowest melting point?

- (a) Li (b) Na
(c) K (d) Rb

Reason: Rubidium has the largest atomic size among the given metals, resulting in the lowest melting point.

viii. Which ionic compound has the highest melting point?

- (a) NaCl (b) KCl
(c) LiCl (d) RbCl

Reason: Lithium chloride has the strongest ionic bonds due to the small size of the lithium ion, giving it the highest melting point.

ix. Which compound contains both covalent and ionic bonds?

- (a) MgCl_2 (b) NH_4Cl
(c) CaO (d) PCl_3

Reason: Ammonium chloride contains both ionic bonds (between NH_4^+ and Cl^-) and covalent bonds (within the NH_4^+ ion).

x. Which among the following has a double covalent bond?

- (a) Ethane (b) Methane
(c) Ethylene (d) Acetylene

Reason: Ethylene (C_2H_4) contains a double covalent bond between the two carbon atoms.

ANSWERS

(i)	(c)	(ii)	(a)	(iii)	(b)	(iv)	(d)	(v)	(c)
(vi)	(a)	(vii)	(d)	(viii)	(c)	(ix)	(b)	(x)	(c)

Additional MCQ's Question and Answers

1. Which type of bond is present in water (H_2O)?

- (a) Ionic bond (b) Non-polar covalent bond
(c) Polar covalent bond (d) Metallic bond

Reason: The bond between hydrogen and oxygen in water is polar due to the electronegativity difference.

2. What type of forces are responsible for the high boiling point of water?

- (a) Dipole-dipole forces (b) Van der Waals forces
(c) Hydrogen bonding (d) Ionic forces

Reason: Hydrogen bonds between water molecules make its boiling point higher.

3. Which bond involves the transfer of electrons?

- (a) Covalent bond (b) Ionic bond
(c) Metallic bond (d) Hydrogen bond

Reason: Ionic bonds are formed by the transfer of electrons from one atom to another.

4. What is the physical state of ionic compounds at room temperature?

- (a) Gas (b) Liquid
(c) Solid (d) Plasma

Reason: Ionic compounds form crystalline solids at room temperature due to strong electrostatic forces.

5. Which of the following has the strongest intermolecular forces?

- (a) NH_3 (b) CH_4
(c) HCl (d) HF

Reason: HF has strong hydrogen bonding, making its intermolecular forces stronger than the others.

6. Which property is most characteristic of ionic compounds?

- (a) High electrical conductivity in solid state (b) High melting and boiling points
(c) Insolubility in water (d) Soft and malleable

Reason: Ionic compounds have strong electrostatic forces, requiring more energy to break.

7. Which of the following is a non-polar covalent compound?

- (a) NaCl (b) H_2
(c) H_2O (d) NH_3

Reason: H_2 has a non-polar covalent bond as both hydrogen atoms have equal electronegativity.

8. Which form of carbon is the hardest naturally occurring substance?

- (a) Graphite (b) Diamond
(c) Coal (d) Charcoal

Reason: Diamond has a rigid structure due to strong covalent bonds between carbon atoms.

9. What type of bond exists in graphite?

- (a) Ionic bond (b) Metallic bond
(c) Covalent bond with delocalized electrons (d) Hydrogen bond

Reason: Graphite has covalent bonds in layers and delocalized electrons, making it conductive.

10. Which compound exists as a liquid at room temperature?

- (a) Methane (CH_4) (b) Hydrogen sulfide (H_2S)
(c) Water (H_2O) (d) Carbon dioxide (CO_2)

Reason: Water is liquid at room temperature due to strong hydrogen bonding.

11. Which property of metals is explained by metallic bonding?

- (a) Low density (b) High melting point
(c) Electrical conductivity (d) Brittle nature

Reason: The sea of delocalized electrons in metallic bonds allows metals to conduct electricity.

12. Why are ionic compounds brittle?

- (a) Due to low melting points (b) Due to weak bonds
(c) Due to their rigid crystal lattice structure (d) Due to mobile electrons

Reason: Applying force disturbs the lattice, causing ionic compounds to break.

13. Which of the following has both covalent and ionic bonds?

- (a) CH_4 (b) H_2O
(c) NaOH (d) NH_3

Reason: NaOH has ionic bonds between Na^+ and OH^- and a covalent bond within the OH^- group.

14. Which compound has the highest electronegativity difference between its elements?

- (a) HCl (b) NaCl
(c) H_2O (d) CO_2

Reason: Sodium and chlorine have the greatest difference in electronegativity, forming a strong ionic bond.

15. Why does graphite conduct electricity while diamond does not?

- (a) Graphite has free ions (b) Graphite has delocalized electrons between layers
(c) Graphite has weak covalent bonds (d) Graphite has a metallic structure

Reason: The delocalized electrons in graphite allow it to conduct electricity, unlike diamond.

ANSWERS

1.	(c)	2.	(c)	3.	(b)	4.	(c)	5.	(d)
6.	(b)	7.	(b)	8.	(b)	9.	(c)	10.	(c)
11.	(c)	12.	(c)	13.	(c)	14.	(b)	15.	(b)

Questions for Short Answers

i. **What type of elements lose their outer electrons easily and what type of elements gain electrons easily?**

Ans. Metals lose electrons easily, while non-metals gain electrons easily.

ii. **Why do lower molecular mass covalent compounds exist as gases or low boiling liquids?**

Ans. They have weak intermolecular forces of attraction, requiring less energy to break.

iii. **Give one example of an element which exists as a crystalline solid and has covalent bonds between its atoms.**

Ans. Diamond is a crystalline solid with covalent bonds between carbon atoms.

iv. **Which property of metals makes them malleable and ductile?**

Ans. The presence of a sea of delocalized electrons allows metal atoms to slide past each other without breaking.

v. **Is coordinate covalent bond a strong bond?**

Ans. Yes, it is as strong as a regular covalent bond.

vi. **Write down the dot and cross formula of HNO_3 .**

Ans. Here is the dot and cross representation of HNO_3 :

(i) Central atom: Nitrogen (N)

(ii) Double bond between N and one oxygen (O); with two lone pairs on oxygen.

(iii) Single bonds between N and the other two oxygens.

- (iv) One oxygen has an H atom attached (O-H).
- (v) The oxygen in the single bond carries a negative charge, balanced by the positive charge on H.

Additional Short Question and Answers

- 1. Why do metals lose electrons easily?**
Ans. Metals lose electrons easily because they have low ionization energy and a tendency to achieve a stable configuration.
- 2. What type of bonding is present in NaCl?**
Ans. NaCl has ionic bonding, formed by the transfer of an electron from sodium to chlorine.
- 3. Why are covalent bonds directional?**
Ans. Covalent bonds are directional because shared electrons are localized between specific atoms.
- 4. What is the primary reason for the hardness of diamond?**
Ans. Diamond's hardness is due to its rigid tetrahedral covalent bonding structure.
- 5. What property of water makes it a good solvent?**
Ans. Water is a polar molecule, which allows it to dissolve ionic and polar compounds effectively.
- 6. Why are ionic compounds soluble in water?**
Ans. Water breaks the electrostatic forces between ions and hydrates them.
- 7. What type of bond exists in O₂?**
Ans. O₂ contains a double covalent bond, where two pairs of electrons are shared.
- 8. Why do molecular compounds like CH₄ have low boiling points?**
Ans. CH₄ has weak van der Waals forces between its molecules, requiring less energy to break.
- 9. What is hydrogen bonding?**
Ans. It is a strong dipole-dipole attraction between hydrogen and highly electronegative atoms like F, O, or N.
- 10. Why are metals good conductors of electricity?**
Ans. Metals have delocalized electrons that move freely, allowing current to flow.
- 11. Which form of carbon is used as a lubricant and why?**
Ans. Graphite, because its layers can slide past each other easily due to weak interlayer forces.
- 12. What makes ionic solids brittle?**
Ans. External forces misalign ions, creating repulsive forces that cause the crystal to break.
- 13. What determines the strength of intermolecular forces?**
Ans. The polarity and molecular size of the substance determine the strength of intermolecular forces.
- 14. Why is F₂ a gas while I₂ is a solid at room temperature?**
Ans. F₂ has weaker intermolecular forces compared to I₂, which has stronger van der Waals forces.
- 15. Why do ionic compounds conduct electricity in molten or aqueous states?**
Ans. In these states, ions are free to move, allowing the conduction of electricity.

Constructed Response Questions

- i. **Why is HF a liquid while HCl is a gas?**
Ans. HF forms strong hydrogen bonds, which increase intermolecular attraction, while HCl only has weaker dipole-dipole forces.
- ii. **Why are covalent compounds generally not soluble in water?**
Ans. Covalent compounds are usually non-polar and cannot interact effectively with polar water molecules.
- iii. **How do metals conduct heat?**
Ans. Metals conduct heat through the free movement of delocalized electrons and lattice vibrations.
- iv. **How many oxides does nitrogen form? Write down the formulae of oxides.**
Ans. Nitrogen forms five oxides: N_2O , NO , NO_2 , N_2O_3 , and N_2O_5 .
- v. **What will happen if NaBr is treated with $AgNO_3$ in water?**
Ans. A white precipitate of $AgBr$ will form due to the reaction between Ag^+ and Br^- ions.
- vi. **Why does iodine exist as a solid while Cl_2 exists as a gas?**
Ans. Iodine has stronger van der Waals forces due to its larger molecular size compared to chlorine.

Descriptive Questions

- i. **Explain the formation of an ionic bond and a covalent bond.**
Ans. See Q. No.2
- ii. **How do ions arrange themselves to form NaCl crystal.**
Ans. **Arrangement of Ions in NaCl Crystal**
- Crystal Structure**
In sodium chloride ($NaCl$), ions arrange themselves in a three-dimensional cubic lattice structure.
 - Ionic Bonding**
The arrangement is driven by strong electrostatic forces of attraction between positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-).
 - Coordination Number**
Each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions. This gives the structure a coordination number of six.
 - Stability**
This alternating pattern of oppositely charged ions ensures maximum stability and a rigid, crystalline structure.
 - Shape**
The overall shape of the $NaCl$ crystal is cubic, reflecting the regular and symmetrical arrangement of the ions.
- iii. **Explain the properties of metals keeping in view the nature of metallic bond.**
Ans. See Q. No. 10
- iv. **Compare the properties of ionic and covalent compounds.**
Ans. See Q. No.13

v. How will you explain the electrical conductivity of graphite crystals?

Ans. See Q. No.

vi. Why are metals usually hard and heavy?

Ans. See Q. No.10

Investigative Questions

i. The formula of $AlCl_3$ in vapour phase is Al_2Cl_6 which means it exists as a dimer. Explain the bonding between its two molecules?

Ans. **Bonding in Al_2Cl_6 Molecule**

1. **Dimerization of $AlCl_3$**

In the vapor phase, aluminum chloride ($AlCl_3$) exists as a dimer, Al_2Cl_6 . This is because $AlCl_3$ molecules are electron-deficient and tend to stabilize by forming coordinate covalent bonds.

2. **Formation of Coordinate Covalent Bonds**

In Al_2Cl_6 , two $AlCl_3$ molecules are linked through two chlorine atoms. These chlorine atoms act as bridges by donating a lone pair of electrons to the aluminum atom of the adjacent molecule, forming coordinate covalent bonds.

3. **Structure of Al_2Cl_6**

The structure of Al_2Cl_6 consists of two aluminum atoms bonded to three terminal chlorine atoms each and two shared chlorine atoms. These shared chlorine atoms create the bridge between the two $AlCl_3$ units.

4. **Stability**

This dimeric structure minimizes electron deficiency on the aluminum atoms, making the molecule more stable in the vapor phase.

ii. **Explain the structure of sand (SiO_2).**

Structure of Sand (SiO_2)

1. **Molecular Structure**

Silicon dioxide (SiO_2), commonly known as sand, has a three-dimensional network structure. In this structure, each silicon atom is covalently bonded to four oxygen atoms in a tetrahedral arrangement. Each oxygen atom is shared between two silicon atoms, forming a continuous network of Si - O bonds.

2. **Tetrahedral Bonding**

In SiO_2 , each silicon atom is at the center of a tetrahedron formed by four oxygen atoms. The oxygen atoms, in turn, are bonded to two silicon atoms. This bonding creates a strong, rigid, and extensive network structure throughout the material.

3. **Crystal Lattice**

SiO_2 forms a crystalline solid structure in its pure form (such as quartz). This crystal lattice extends in all directions, making the material hard and giving it high melting and boiling points. The lattice structure is highly stable and contributes to the strength and durability of sand.

4. **Properties of Sand**

Due to the strong covalent bonds between silicon and oxygen, SiO_2 is chemically inert, resistant to weathering, and insoluble in water. The structure of SiO_2 also explains its transparency in its most common forms (like quartz crystals) and its use in various industries like glass-making.

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