Chapter 5:
Chemical Bonding

Short Questions (Exercise)

(i) State octet and duplet rules.

Octet Rule: Atoms gain, lose, or share electrons to achieve 8 electrons in their valence shell, resembling a noble gas configuration.

Duplet Rule: Smaller atoms like Hydrogen and Helium aim to achieve 2 electrons in their valence shell for stability.

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

Two nitrogen atoms share three pairs of electrons (a triple bond) to complete their octet.

The structure is: N ≡ N

Each nitrogen atom shares three electrons, resulting in a stable configuration.

(iii) How does Al form a cation?

Aluminium (Al) loses 3 electrons from its valence shell ($3s^2\ 3p^1$) to form a Al^{3+} cation:

$Al \rightarrow Al^{3+} + 3e^-$

(iv) How does O form an anion?

Oxygen (O) gains 2 electrons to complete its octet, forming O^{2-}:

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

(v) Draw the electron cross and dot structure for H_2O:

Water (H_2O) molecule structure:

H : O : H

Oxygen shares one electron with each hydrogen atom. Oxygen's octet is complete, and each hydrogen satisfies the duplet rule.

3. Describe the importance of noble gas electronic configuration.

Noble gas electronic configurations represent a stable arrangement of electrons. Atoms achieve this configuration by gaining, losing, or sharing electrons, ensuring chemical stability and low reactivity.

4. Explain how elements attain stability.

Elements attain stability by:

1. Gaining electrons (e.g., O^{2-}).
2. Losing electrons (e.g., Na^{+}).
3. Sharing electrons (e.g., H_2).

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

Bonds are formed by:

1. Ionic bonding: Transfer of electrons between atoms (e.g., NaCl).
2. Covalent bonding: Sharing of electrons between atoms (e.g., H_2).
3. Metallic bonding: Free electrons shared within a lattice of metal ions.

6. Formation of covalent bond between two non-metallic elements.

A covalent bond is formed when two non-metallic elements share electrons. For example, in H_2O, oxygen shares one electron with each hydrogen atom.

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

Single Bond: Sharing of one pair of electrons (e.g., H_2).
Double Bond: Sharing of two pairs of electrons (e.g., O_2).
Triple Bond: Sharing of three pairs of electrons (e.g., N_2).

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

a. Boron: 3 (Group IIIA).
b. Neon: 8 (Group VIIIA).
c. Rubidium: 1 (Group IA).
d. Barium: 2 (Group IIA).
e. Arsenic: 5 (Group VA).

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

a. Aluminium (Al^{3+}): [Al] loses $3e^{-} \rightarrow Al^{3+}$.
b. Strontium (Sr^{2+}): [Sr] loses $2e^{-} \rightarrow Sr^{2+}$.
c. Barium (Ba^{2+}): [Ba] loses $2e^{-} \rightarrow Ba^{2+}$.

10. Sulphur isotope questions:

a. Define the term isotope:
Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.

b. Define relative atomic mass:
Relative atomic mass is the weighted average of the masses of the isotopes of an element compared to C-12.

c. Calculate the relative atomic mass of sulphur:
Relative Atomic Mass = ((32 × 95) + (33 × 0.76) + (34 × 4.22)) / 100 = 32.09.

d. Complete the table:

Isotope	Protons	Neutrons	Electrons
S-32	16	16	16
S-34	16	18	16

e. Where will you place S in the periodic table?
Group: VIA (16), Period: 3.

f. How many electrons will S gain to acquire stability?
Sulphur will gain 2 electrons to form S^{2-}.

g. How many atoms of S are there in 0.3 moles of sulphur?
Atoms of S = $0.3 \times 6.022 \times 10^{23} = 1.807 \times 10^{23}$.

11. Questions about an atom with atomic number 9 and mass number 19:

a. State the number of protons and neutrons:
Protons: 9, Neutrons: 19 - 9 = 10.

b. State the number of electrons:
Electrons: 9.

c. Show the electron cross-dot diagram:
$[F]^-$ (fluoride ion).

d. Write electronic configuration:
$1s^2\ 2s^2\ 2p^5$.

e. Point out its group in the periodic table:
Group VIIA (17).

f. Point out its period in the periodic table:
Period 2.

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

(a) What is the charge on these ions?

- Magnesium ion (Mg^{2+}): +2
- Oxide ion (O^{2-}): -2

(b) How these ions get these charges?

- Magnesium loses 2 electrons from its outermost shell to form Mg^{2+}.
- Oxygen gains 2 electrons to complete its octet and forms O^{2-}.

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

Magnesium transfers 2 electrons to oxygen.
The resulting structure is: Mg^{2+} and O^{2-}.

13. Draw structures of the ions formed when these atoms react (Calcium and Oxygen).

- Calcium loses 2 electrons, forming Ca^{2+}.
- Oxygen gains 2 electrons, forming O^{2-}.
- Resulting compound: CaO.

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

(a) Which substance is a metal?

Substance D: High melting point, conducts electricity in both solid and molten states.

(b) Which substance is an ionic compound?

Substance B: High melting point, conducts electricity in molten state only.

(c) Which substance is a covalent compound?

Substance C: Low melting point, does not conduct electricity.

(d) Which substance is a non-metal?

Substance A: High melting point, does not conduct electricity.

Project

Type of Bond	Definition	Formation Mechanism	Example
Ionic Bond	A bond formed by the transfer of electrons from one atom to another.	Occurs between metals and non-metals. The metal loses electrons (cation), and the non-metal gains electrons (anion).	NaClNaClNaCl (Sodium Chloride): Sodium donates an electron to Chlorine.
Covalent Bond	A bond formed by the sharing of electrons between two atoms.	Typically occurs between two non-metals. Electrons are shared to achieve a stable octet or duplet configuration.	H2OH_2OH2O (Water): Oxygen shares electrons with Hydrogen.
Metallic Bond	A bond formed by the attraction between a lattice of positive metal ions and a sea of delocalized electrons.	Electrons move freely throughout the structure, leading to properties like conductivity and malleability.	Copper (CuCuCu), Iron (FeFeFe).
Hydrogen Bond	A weak bond formed between a hydrogen atom covalently bonded to an electronegative atom and another electronegative atom.	Involves hydrogen and elements like OOO, NNN, or FFF.	H2OH_2OH2O molecules: Hydrogen bonds between water molecules.
Van der Waals Force	Weak intermolecular forces of attraction.	Result from temporary dipoles formed in molecules.	Interactions in Cl2Cl_2Cl2 or noble gases like Argon (ArArAr).

Extera Short Questions (Topic Wise)

5.1: Why Do Atoms React?

Why do atoms react?

Atoms react to achieve a stable electronic configuration, usually a full outer shell of electrons (octet rule).

Example: Sodium reacts with chlorine because sodium loses one electron and chlorine gains it, forming NaCl.

What is the driving force behind chemical reactions?

The drive to lower potential energy and increase stability leads atoms to form bonds.

Example: Hydrogen atoms combine to form H_2, a more stable molecule.

How do noble gases relate to atomic reactivity?

Noble gases are unreactive because they already have a full outer shell of electrons.

Example: Helium and neon do not form compounds under normal conditions.

What happens when atoms react?

Atoms form bonds by gaining, losing, or sharing electrons.

Example: Oxygen gains two electrons from two hydrogen atoms to form water (H_2O).

5.2: Chemical Bonds

What is a chemical bond?

A chemical bond is a force of attraction that holds atoms together to form molecules.

Example: In water (H_2O), covalent bonds hold the hydrogen and oxygen atoms together.

What are the main types of chemical bonds?

The main types are ionic, covalent, and metallic bonds.

Example: NaCl has ionic bonds, H_2O has covalent bonds, and copper has metallic bonds.

What determines the type of bond formed?

The difference in electronegativity between atoms determines whether a bond is ionic, covalent, or metallic.

Example: Sodium (low electronegativity) and chlorine (high electronegativity) form an ionic bond.

What is bond energy?

Bond energy is the energy required to break a chemical bond.

Example: The bond energy of the H-H bond in hydrogen gas is 436 kJ/mol.

5.3: Types of Bonds

What is an ionic bond?

An ionic bond forms when one atom transfers electrons to another, creating oppositely charged ions.

Example: Sodium (Na) donates an electron to chlorine (Cl), forming NaCl.

What is a covalent bond?

A covalent bond forms when atoms share electrons to achieve stability.

Example: Two hydrogen atoms share electrons to form H_2.

What is a metallic bond?

A metallic bond is the attraction between metal cations and a "sea" of free-moving electrons.

Example: In copper (Cu), delocalized electrons allow conductivity and malleability.

Which bond is the strongest?

Covalent bonds are generally the strongest due to shared electron pairs.

Example: The covalent bonds in diamonds make them extremely hard.

5.4: Intermolecular Forces

What are intermolecular forces?

Intermolecular forces are weak forces of attraction between molecules.

Example: Water molecules are held together by hydrogen bonds, giving water its high boiling point.

What are the types of intermolecular forces?

London dispersion forces: Weak forces in all molecules.

Dipole-dipole interactions: Attraction between polar molecules.

Hydrogen bonding: Strong dipole-dipole interaction involving hydrogen and N, O, or F.

Example: Hydrogen bonding occurs in water (H_2O).

What is hydrogen bonding?

Hydrogen bonding is a strong attraction between a hydrogen atom and a highly electronegative atom like oxygen, nitrogen, or fluorine.

Example: Hydrogen bonds in DNA hold the two strands of the helix together.

How do intermolecular forces affect boiling points?

Stronger intermolecular forces lead to higher boiling points.

Example: Water (H_2O) has a higher boiling point than methane (CH_4) due to hydrogen bonding.

5.5: Nature of Bonding, Structure, and Properties

How does bonding influence properties?

Bond type affects melting points, solubility, and conductivity.

Example: Ionic compounds like NaCl have high melting points, while covalent compounds like CH_4 have low melting points.

Why are ionic compounds solid at room temperature?

Strong electrostatic forces hold ions together in a lattice structure.

Example: Table salt (NaCl) forms a solid crystalline structure.

Why do covalent compounds have low melting points?

Weak intermolecular forces hold the molecules together.

Example: Methane (CH_4) has a low melting point due to weak London dispersion forces.

Why are metals good conductors?

Metals have free-moving electrons that carry electric current.

Example: Copper is used in electrical wiring due to its high conductivity.

5.6: Metallic Bonds

What is the metallic bond model?

Metallic bonds consist of a lattice of metal cations surrounded by a "sea" of delocalized electrons.

Example: In aluminum, free electrons allow the metal to conduct heat and electricity.

Why are metals malleable and ductile?

Metallic bonds allow metal atoms to slide over one another without breaking.

Example: Gold can be hammered into thin sheets without cracking.

Why do metals conduct electricity?

The free-moving electrons in metallic bonds carry electrical current efficiently.

Example: Silver is the best conductor of electricity due to its metallic bonding.

Which metal property results from metallic bonding?

Metals have high melting points due to strong metallic bonds.

Example: Tungsten has a high melting point, making it suitable for light bulb filaments.

Extera Long Questions (Topic Wise)

1. Why do atoms react, and how do chemical bonds form?

Why Do Atoms React?

- Atoms react to achieve a stable electronic configuration, typically a full outer electron shell, following the **octet rule** (eight electrons in the outer shell).
- This is driven by the need to lower their potential energy and increase stability.

How Do Chemical Bonds Form?

1. **Ionic Bonds:**
    - Formed when one atom transfers electrons to another, creating oppositely charged ions that attract each other.
    - **Example:** Sodium (Na) loses one electron to become Na^+, and chlorine (Cl) gains an electron to become Cl^-, forming NaCl.
2. **Covalent Bonds:**
    - Formed when atoms share electrons to fill their outer shells.
    - **Example:** Two hydrogen atoms share one electron pair to form H_2, stabilizing both atoms.
3. **Metallic Bonds:**
    - Formed when metal atoms share a "sea" of delocalized electrons.
    - **Example:** In copper (Cu), free-moving electrons allow electrical conductivity.

Significance of Reactions

- Chemical reactions allow the formation of compounds with unique properties. For example, water (H_2O) is essential for life and results from the reaction of hydrogen and oxygen.

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2. What are the types of chemical bonds, and how do their properties differ?

Types of Chemical Bonds

1. **Ionic Bonds:**
    - Formed by the transfer of electrons between metals and non-metals.
    - Strong electrostatic forces hold ions together in a crystal lattice.
    - **Properties:** High melting and boiling points, conductivity in molten/aqueous states.
    - **Example:** NaCl (sodium chloride).
2. **Covalent Bonds:**

- Formed when atoms share electron pairs.
- Can be single (H-H), double (O=O), or triple (N≡N).
- **Properties:** Low melting and boiling points, poor conductivity.
- **Example:** H_2O (water).

3. **Metallic Bonds:**
   - Formed by the attraction between metal cations and a sea of delocalized electrons.
   - **Properties:** High conductivity, malleability, and ductility.
   - **Example:** Copper (Cu).

Comparison of Bond Strengths

- Covalent bonds are strong within molecules, while ionic bonds are strong in crystalline solids.
- Metallic bonds provide metals with unique properties like ductility and conductivity.

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3. What are intermolecular forces, and how do they affect the properties of substances?

Definition of Intermolecular Forces

Intermolecular forces are weak attractions between molecules, distinct from strong intramolecular bonds. They play a crucial role in determining the physical properties of substances.

Types of Intermolecular Forces

1. **London Dispersion Forces (LDFs):**
   - Present in all molecules, caused by temporary dipoles.
   - **Example:** Methane (CH_4) has weak dispersion forces, resulting in a low boiling point.
2. **Dipole-Dipole Interactions:**
   - Occur in polar molecules where permanent dipoles attract each other.
   - **Example:** Hydrogen chloride (HCl) exhibits dipole-dipole forces.
3. **Hydrogen Bonding:**
   - A strong dipole-dipole force involving hydrogen bonded to highly electronegative atoms like N, O, or F.
   - **Example:** Water (H_2O) molecules form hydrogen bonds, giving water its high boiling point and surface tension.

Effects on Properties

- **Boiling and Melting Points:** Stronger intermolecular forces lead to higher boiling/melting points.
- **Solubility:** Polar molecules dissolve in polar solvents, while non-polar molecules dissolve in non-polar solvents.

- **Example:** Oil (non-polar) does not dissolve in water (polar).

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4. What are metallic bonds, and how do they influence the properties of metals?

What Are Metallic Bonds?

- Metallic bonds are formed by the attraction between metal cations and a "sea" of delocalized electrons.
- These electrons are free to move throughout the lattice, giving metals their unique properties.

Properties of Metals Due to Metallic Bonding

1. **High Electrical and Thermal Conductivity:**
    - Free-moving electrons transfer energy efficiently.
    - **Example:** Copper is used in electrical wiring.
2. **Malleability and Ductility:**
    - Metal atoms can slide over one another without breaking the bond.
    - **Example:** Gold can be hammered into thin sheets (malleability).
3. **High Melting and Boiling Points:**
    - Strong metallic bonds require significant energy to break.
    - **Example:** Tungsten has one of the highest melting points among metals.
4. **Shiny Appearance (Luster):**
    - Free electrons reflect light, giving metals their shiny appearance.
    - **Example:** Aluminum is used in reflective surfaces.

Applications of Metallic Bonding

- **Construction:** Steel (an alloy of iron) is strong and malleable.
- **Transportation:** Aluminum is lightweight yet durable, making it ideal for aircraft.
- **Electronics:** Gold and silver are used in circuits due to excellent conductivity.

MCQS

1. **Why do atoms react?**
   a) To lose energy
   b) To achieve a stable electronic configuration ☑
   c) To gain mass
   d) To break existing bonds
2. **What rule explains the tendency of atoms to react?**
   a) Aufbau principle
   b) Octet rule ☑
   c) Hund's rule
   d) Periodic law
3. **Which atom is most stable without reacting?**
   a) Neon (Ne) ☑
   b) Sodium (Na)
   c) Oxygen (O)
   d) Hydrogen (H)
4. **What is an ionic bond?**
   a) A bond formed by sharing electrons
   b) A bond formed between two metals
   c) A bond formed by the transfer of electrons ☑
   d) A bond formed by hydrogen atoms
5. **Which compound has an ionic bond?**
   a) H_2O
   b) NaCl ☑
   c) CO_2
   d) O_2
6. **What is a covalent bond?**
   a) A bond formed by losing electrons
   b) A bond formed by sharing electrons ☑
   c) A bond formed between metal atoms
   d) A bond formed by free electrons
7. **Which compound has a covalent bond?**
   a) NaCl
   b) CH_4 ☑
   c) MgO
   d) KBr
8. **What is a metallic bond?**
   a) A bond between non-metals
   b) A bond due to electron sharing
   c) A bond involving a "sea" of delocalized electrons ☑
   d) A bond in polar molecules
9. **Which property is due to metallic bonding?**
   a) Low melting point

b) Poor conductivity

c) Malleability ☑

d) High reactivity

10. **Which is a characteristic of ionic compounds?**

a) High melting points ☑

b) Low solubility in water

c) Poor conductivity in molten state

d) Brittle but conductive solids

11. **Which intermolecular force is the strongest?**

a) London dispersion forces

b) Dipole-dipole interactions

c) Hydrogen bonding ☑

d) Ionic forces

12. **Which molecule exhibits hydrogen bonding?**

a) CH_4

b) CO_2

c) H_2O ☑

d) NaCl

13. **What is the main force in non-polar molecules?**

a) London dispersion forces ☑

b) Hydrogen bonding

c) Ionic bonds

d) Metallic bonds

14. **Why does water have a high boiling point?**

a) It has ionic bonds

b) It forms hydrogen bonds ☑

c) It has low molecular weight

d) It is non-polar

15. **What type of bond forms in NaCl?**

a) Covalent

b) Ionic ☑

c) Metallic

d) Hydrogen

16. **What type of bond forms in H_2O?**

a) Ionic

b) Covalent ☑

c) Metallic

d) London dispersion

17. **Why are metals good conductors of electricity?**

a) Their atoms are tightly packed

b) They have free-moving electrons ☑

c) They have low melting points

d) They contain covalent bonds

18. **Which property is unique to metals due to metallic bonding?**

a) Brittle nature

b) Malleability ☑

c) Low density
d) High electronegativity

19. **What type of bond exists in CO_2?**
a) Ionic
b) Covalent ☑
c) Metallic
d) Hydrogen

20. **Which molecule contains a triple bond?**
a) H_2
b) N_2 ☑
c) O_2
d) HCl

21. **What is a polar covalent bond?**
a) A bond with unequal sharing of electrons ☑
b) A bond with equal sharing of electrons
c) A bond with no electron sharing
d) A bond between ions

22. **What property results from strong intermolecular forces?**
a) Low boiling point
b) High boiling point ☑
c) Low viscosity
d) High volatility

23. **Which bond is the weakest?**
a) Covalent bond
b) Ionic bond
c) Hydrogen bond
d) London dispersion forces ☑

24. **What causes hydrogen bonding?**
a) Electrons in the outer shell
b) Attraction between hydrogen and N, O, or F ☑
c) Free-moving electrons
d) Metallic character

25. **Which is an example of a compound with metallic bonding?**
a) H_2O
b) NaCl
c) O_2
d) Fe (Iron) ☑

26. **Why do ionic compounds conduct electricity in solution?**
a) They are polar molecules
b) Ions are free to move ☑
c) They form covalent bonds
d) They have metallic properties

27. **What happens to the electrons in a metallic bond?**
a) They move freely between atoms ☑
b) They are shared between specific atoms
c) They are transferred to non-metals
d) They stay fixed in place

28. **Which type of bond holds the two strands of DNA together?**
    a) Covalent bonds
    b) Hydrogen bonds ☑
    c) Ionic bonds
    d) Metallic bonds
29. **What determines the type of bond between two atoms?**
    a) The number of neutrons
    b) The number of valence electrons
    c) The electronegativity difference ☑
    d) The atomic mass
30. **Which property makes metals useful for electrical wiring?**
    a) High melting points
    b) High electrical conductivity ☑
    c) High density
    d) High reactivity