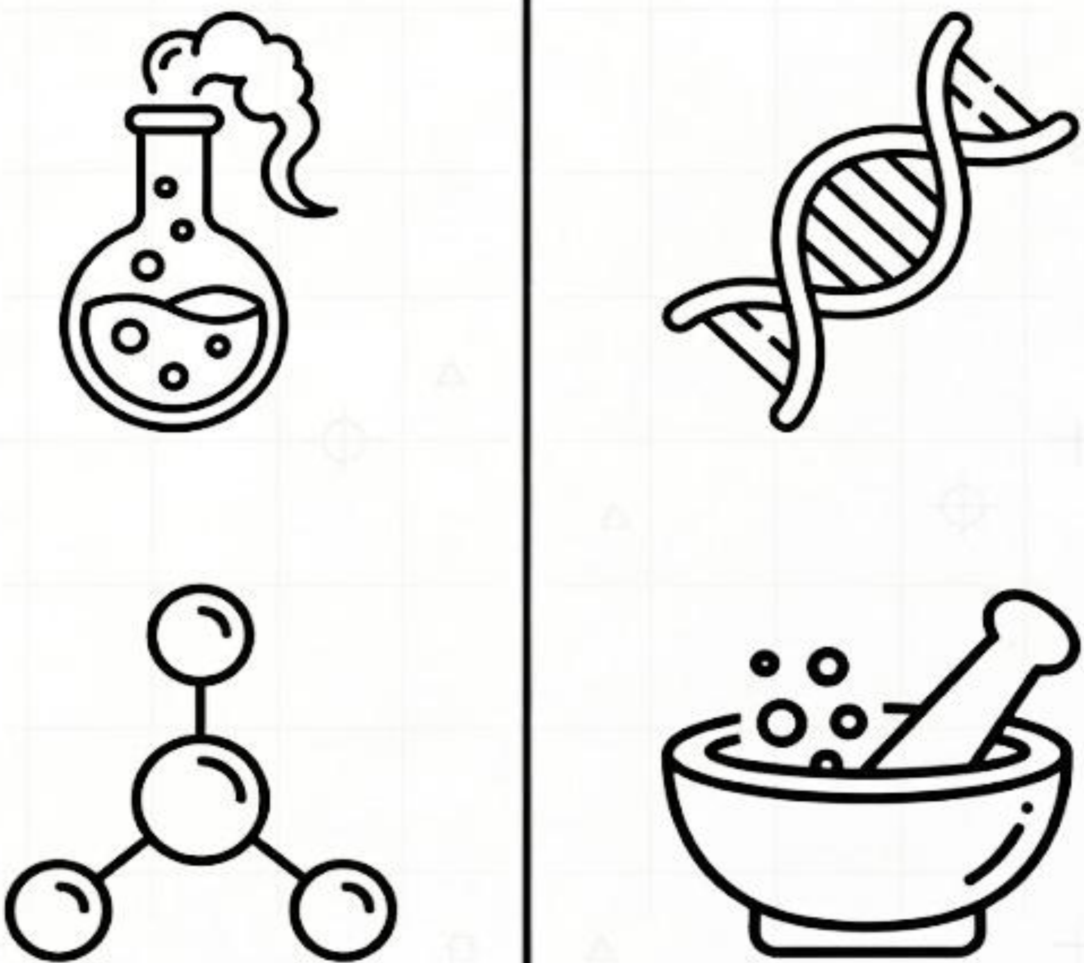
MVS FOUNDATION

FREE NOTES

CHEMISTRY (313)

CHAPTERWISE NOTES



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CHEMISTRY

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Component	Details	Marks
Public Exam (Selected Modules 1,2,3,4,5,8)	Total Chapters : 8	44
Practical Exam	Practical	20
TMA	Tutor Marked Assignment	16
Final Possible Marks		**80 Marks**

1

Atoms, Molecules and Chemical Arithmetic

Chemistry's Role: Essential in health (gene therapy, antibiotics), energy (solar/nuclear fusion), and environment (reducing greenhouse gases).

Particulate Nature: Matter consists of tiny particles called atoms (from Greek 'atmos' meaning indivisible).

1. Laws of Chemical Combination

1. **Law of Conservation of Mass:** Total mass of reactants equals the total mass of products in a chemical reaction.
2. **Law of Constant Proportions:** A pure chemical compound always contains its component elements in a fixed ratio by mass.
3. **Law of Multiple Proportions:** When two elements form multiple compounds, the masses of one element that combine with a fixed mass of the other are in a ratio of small whole numbers.

2. Dalton's Atomic Theory

Matter is composed of indivisible atoms.

Atoms of the same element are identical in mass and properties.

Atoms are indestructible and retain identity during chemical reactions.

3. Measurement and SI Units

Measurement is essential in every walk of life and requires a "**unit**" or "**reference standard**". In **1960**, the 'International System of Units' (**SI units**) was proposed to provide a universal metric system.

1. The Seven SI Base Units

These seven units correspond to the seven base physical quantities:

Physical Quantity	Name of SI Unit	Symbol
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s

Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

2. SI Prefixes for Multiples and Sub-multiples

Prefixes are used to denote very large or very small quantities. These symbols are prefixed to the unit symbol.

Prefix	Symbol	Meaning	Example
Tera	T	10^{12}	1 terametre (Tm) = 1.0×10^{12} m
Giga	G	10^{9}	1 gigametre (Gm) = 1.0×10^{9} m
Mega	M	10^{6}	1 megametre (Mm) = 1.0×10^{6} m
Kilo	k	10^{3}	1 kilometre (km) = 1.0×10^{3} m
Hecta	h	10^{2}	1 hectametre (hm) = 1.0×10^{2} m
Deca	da	10^{1}	1 decametre (dam) = 1.0×10^{1} m
Deci	d	10^{-1}	1 decimetre (dm) = 1.0×10^{-1} m
Centi	c	10^{-2}	1 centimetre (cm) = 1.0×10^{-2} m
Milli	m	10^{-3}	1 millimetre (mm) = 1.0×10^{-3} m
Micro	μ	10^{-6}	1 micrometre (μm) = 1.0×10^{-6} m
Nano	n	10^{-9}	1 nanometre (nm) = 1.0×10^{-9} m
Pico	p	10^{-12}	1 picometre (pm) = 1.0×10^{-12} m

4. Mole Concept and Stoichiometry

Mole (mol): Amount of substance containing **6.022×10^{23}** (Avogadro's Constant) entities.

Molar Mass: Mass of one mole of a substance in grams (**g/mol**)

Molar Volume: One mole of any ideal gas occupies **22.7 L** at **STP (273 K, 1 bar).**

Empirical Formula: Simplest whole-number ratio of atoms in a compound.

Molecular Formula: Actual number of atoms of each element in a molecule.

Limiting Reagent: The reactant that is completely consumed first, limiting the product formed.

1. Visual Example: Limiting Reagent

The **Limiting Reagent** is the substance that is used up completely and stops the reaction.

Reaction: $2H_2 + O_2 \rightarrow 2H_2O$

Before Reaction	After Reaction
4 molecules of H_2	0 molecules of H_2 (Used up)
3 molecules of O_2	1 molecule of O_2 (Left over)
	4 molecules of H_2O (Formed)

Conclusion: In this case, **H_2 is the Limiting Reagent** because it ran out first.

2. Visual Example: Empirical Formula Calculation

The **Empirical Formula** is the simplest whole-number ratio of atoms.

Example: A compound has **53.1% Carbon** and **46.9% Oxygen**.

Element	% Mass	Moles (% /At. Mass)	Simple Ratio (÷ Smallest)	Result
Carbon (C)	53.1%	53.1 / 12 = 4.43	4.43/ 2.93 =1.5	3
Oxygen (O)	46.9%	46.9 / 16 = 2.93	2.93 / 2.93 =1	2

Empirical Formula: C_3O_2

Top 10 Most Expected Question of this Year with Solutions

Question 1: State the Law of Constant Proportions with an example.

Solution: This law states that in a given chemical compound, the proportions by mass of the constituent elements are always fixed.

Example: In pure water **(H_2O),** the ratio of the mass of hydrogen to oxygen is always **1:8**, regardless of whether the water is obtained from a well, a river, or synthesized in a lab.

Question 2: Calculate the molar mass of Sodium Carbonate ($Na_2CO_3 \circ 10H_2O$).

Solution: Molar mass = (2 x Mass of Na) + Mass of C + (3 x Mass of O) + 10 x (Mass of H_2O)

Molar mass = (2 x 23) + 12 + (3 x 16) + 10 x (2 x 1 + 16)

Molar mass = 46 + 12 +48 + 180 = **286** g/mol

Question 3: A sample of nitrogen gas consists of 4.22×10^{23} molecules. How many moles are there?

Solution: $\text{Number of moles} = \frac{\text{Given molecules}}{\text{Avogadro's Constant}}$

- $n = \frac{4.22\times10^{23}}{6.022\times10^{23}} \approx \mathbf{0.70}\ \text{mol.}$

Question 4: What is the empirical formula of Benzene (C_6H_6) and Glucose ($C_6H_{12}O_6$)?

Solution: **Benzene (C_6H_6):** Simplest ratio of **C:H is 1:1.** Empirical Formula = **CH**

Glucose ($C_6H_{12}O_6$): Simplest ratio of **C:H:O is 1:2:1.** Empirical Formula = **CH_2O**

Question 5: Define 'Limiting Reagent'. Why is it significant in chemical calculations?

Solution: The limiting reagent is the reactant that is entirely consumed in a reaction before others. It is significant because it determines the maximum amount of product that can be formed and helps in calculating the excess amount of other reactants left over.

Question 6: Calculate the volume occupied by 2.5 moles of CO_2 gas at STP.

Solution: At **STP, 1** mole of gas **= 22.7 L**

- Volume of **2.5** moles **= 2.5 x 22.7 = 56.75 L**

Question 7: Calculate the percentage of Iron (Fe) in Fe_3O_4. (Atomic mass: Fe=56, O=16)

Solution: Molar mass of **Fe_3O_4 = (3 x 56) + (4 x 16) = 168 + 64 = 232 g/mol**

$$\text{Percentage of Fe} = \left(\frac{\text{Mass of Fe in compound}}{\text{Molar mass}}\right) \times 100$$

$$\%\text{Fe} = \left(\frac{168}{232}\right) \times 100 \approx \mathbf{72.41\%}.$$

Question 8: State any three postulates of Dalton's Atomic Theory.

Solution:

1. Matter consists of tiny, indivisible particles called atoms.
2. All atoms of a specific element are identical in mass and chemical properties.
3. Atoms combine in small whole-number ratios to form compounds.

Question 9: Find the mass of 1.0 x 10^{19} carbon-12 atoms.

Solution: Mass of 6.022 x 10^{19} atom = 12 g

$$\text{Mass of } 1.0 \times 10^{19} \text{ atoms} = \frac{12 \times 1.0 \times 10^{19}}{6.022 \times 10^{23}} = \mathbf{1.99 \times 10^{-4}} \text{ g}.$$

Question 10: How much hydrogen is required to produce 1 metric ton (10^6 g) of Ammonia (NH_3)?

Solution: **Reaction:** $N_2 + 3H_2 \rightarrow 2NH_3$

Mass relationship: 34 g NH_3 requires 6 g H_2.

$$\text{For } 10^6 \text{ g } NH_3,\ H_2 \text{ required} = \frac{6 \times 10^6}{34} = \mathbf{1.76 \times 10^5} \text{ g}.$$

2

Atomic Structure

1. Discovery of Subatomic Particles (The Beginning)

Cathode Rays (Electrons): Discovered by **J.J. Thomson**; these are negatively charged particles that travel from cathode to anode in a discharge tube.

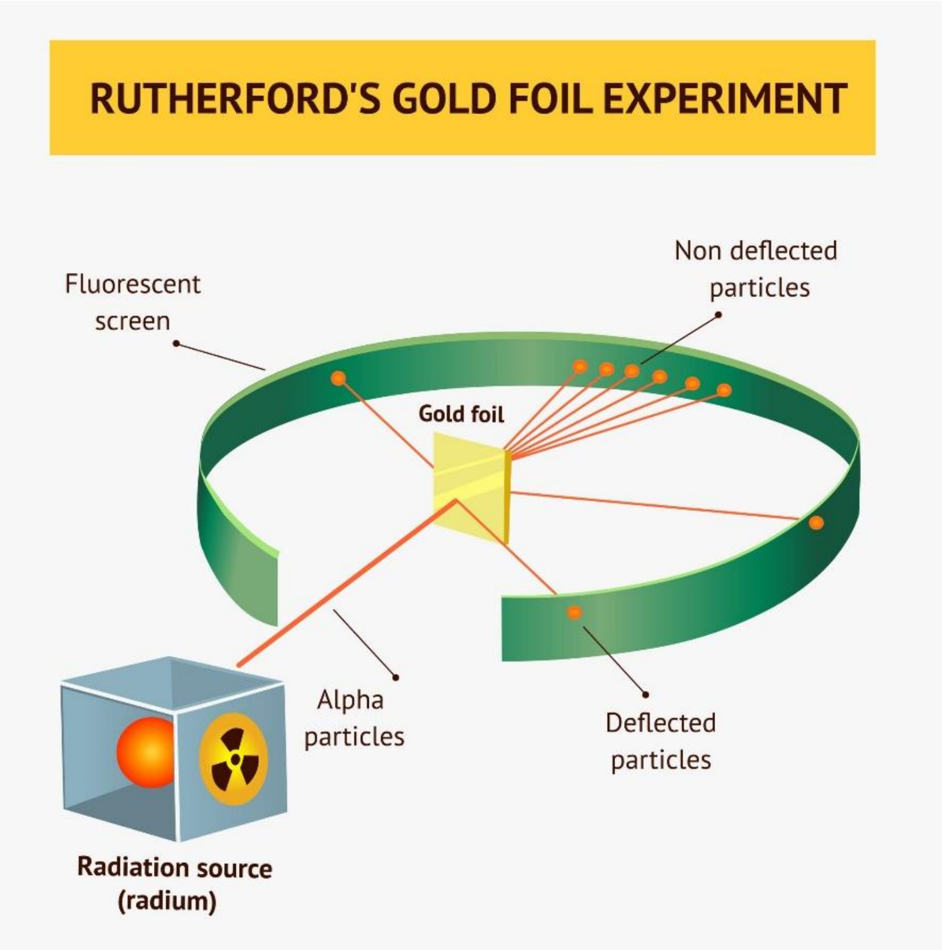
Anode Rays (Protons): Discovered by Goldstein; these are positively charged "canal rays" formed from gas ionization.

Neutrons: Discovered by James Chadwick (**1932**); neutral particles found in the nucleus.

2. Early Atomic Models

Thomson's Model: Positive charge is spread like a pudding with electrons embedded like plums.

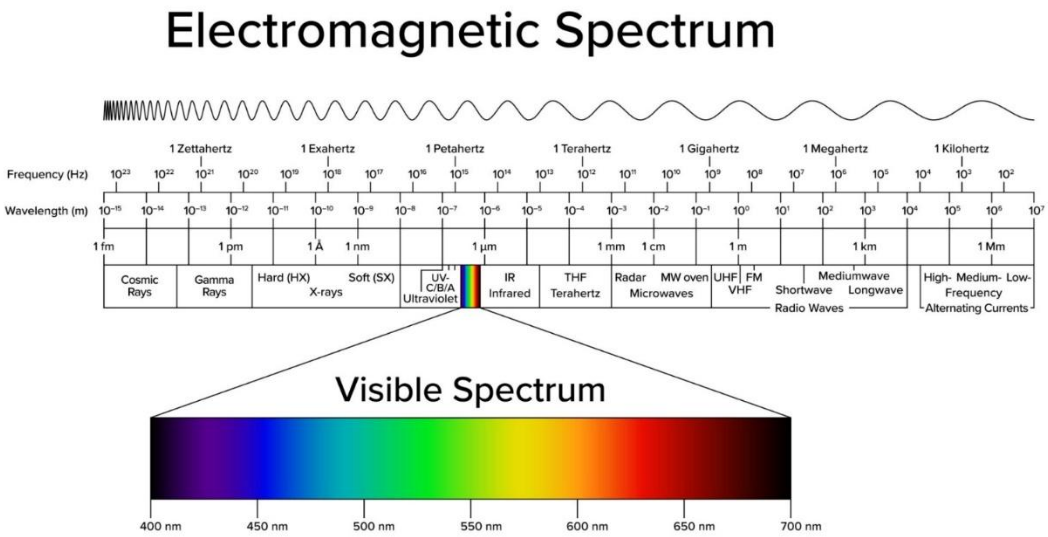
Rutherford's Model: Based on the **α -particle** scattering experiment; discovered the Nucleus and concluded that most of the atom is empty space.



3. Electromagnetic Spectrum

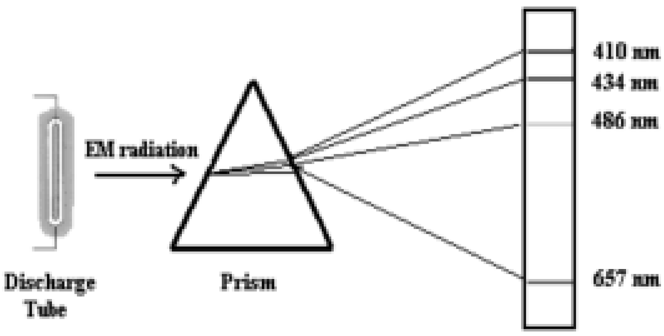
Electromagnetic Radiation and Spectrum Electromagnetic radiations travel in the form of waves with a speed equal to the speed of light **(c).**Order (Increasing Wavelength) :

Gamma rays < X-rays < Ultraviolet (UV) < Visible Light < Infrared (IR) < Microwaves < Radio waves.



4. Line Spectrum of Hydrogen Atom

When this light is passed through a prism it splits up into a set of five lines. This spectrum is called the line spectrum of hydrogen.



$$\bar{\nu} = \frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{cm}^{-1} \; ; \; R_H = 109677 \text{ cm}^{-1}$$

5. Atomic & Mass Numbers

Atomic Number (Z): Total number of protons in the nucleus.

Mass Number (A): Sum of protons and neutrons **(A = Z + n).**

Isotopes: Same **Z,** different **A** (e.g., Protium and Deuterium).

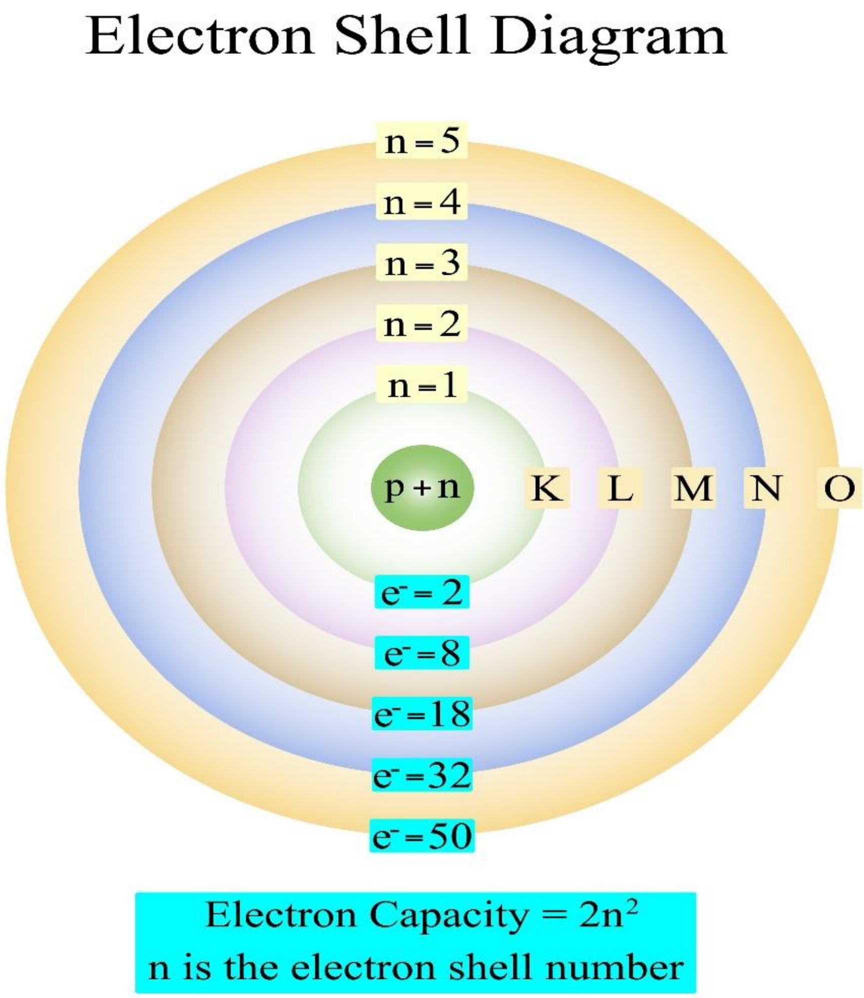
Isobars: Same **A,** different **Z** (e.g., Argon and Calcium).

6. Bohr's Theory & Quantum Nature

Electromagnetic Radiation: Travels as waves where **c = v λ**

Planck's Theory: Energy is absorbed/emitted in packets called Quanta **(E = h\v).**

Bohr's Model: Electrons revolve in fixed circular orbits with quantized energy levels.



Dual Nature (de Broglie): Matter has both wave and particle properties $(\lambda = \frac{h}{mv})$.

Heisenberg Uncertainty Principle: Impossible to know exact position and momentum simultaneously $(\Delta x \cdot \Delta p \geq \frac{h}{4\pi})$.

7. Quantum Numbers

Principal (n): Defines the main shell and energy.

Azimuthal (l): Defines the subshell shape **(s, p, d, f).**

Magnetic (ml): Defines orbital orientation.

Spin (s): Defines the direction of electron spin.

8. Rules for Filling Orbitals (The Final Concept)

1. Aufbau Principle

Electrons occupy the **lowest energy orbital** first.

Order: 1s, 2s, 2p, 3s, 3p, 4s, 3d...

2. Pauli's Exclusion Principle

No two electrons in an atom can have the same 4 quantum numbers.

An orbital can hold **maximum 2 electrons** with opposite spins.

3. Hund's Rule of Maximum Multiplicity

Pairing in degenerate orbitals **(p, d, f)** doesn't start until **each orbital is singly occupied**.

Top 10 PYQs (Solutions)

1. **Calculate p, n, e in $^{80}_{35}Br$.**

Ans: **p = 35, e = 35, n = 80 – 35 = 45**

2. **Explain Heisenberg Uncertainty Principle.**

Ans: Position and momentum of a subatomic particle cannot be measured exactly at the same time.

3. **What are the values of l for n=2?**

Ans: **L = 0 (2s) and L = 1 (2p)**

4. **Write the electronic configuration of Cr (Z=24)**

Ans: **[Ar] $4s^1$ $3d^5$** (Extra stability of half-filled subshell).

5. **Calculate the energy of a photon with frequency 5×10^{14} Hz**

Ans: **E = hv = $6.626 \times 10^{-34} \times 5 \times 10^{14}$ = 3.313 z 10^{-19} J**

6. Calculate the wavelength of the Balmer line corresponding to $n_2 = 3$

Ans: Use the Rydberg formula for the Balmer series (transition to $n_1 = 2$):

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = R\left(\frac{1}{4} - \frac{1}{9}\right) = R \cdot \frac{5}{36}$$
$$\lambda = \frac{36}{5R}$$

With $R \approx 1.097 \times 10^7\ \text{m}^{-1}$:

$$\lambda \approx 6.56 \times 10^{-7}\ \text{m} = 656\ \text{nm}$$

Final Answer: 656 nm (red line, H-alpha)

7. What is Pauli's exclusion principle?

Ans: **Pauli's Exclusion Principle** (proposed by Wolfgang Pauli):

No two electrons in the same atom can have the **same set of four quantum numbers**.

In simple terms:

- An orbital can hold **at most 2 electrons**
- These two electrons must have **opposite spins**

8. What is Aufbau principle? What are (n + l) rules?

Ans: **Aufbau Principle**

The **Aufbau principle** (German for "building up") states that:

Electrons fill atomic orbitals starting from the **lowest energy level first**, then move to higher energy levels.

This determines the **order of electron configuration** in atoms.

(n + l) Rule

The **(n + l) rule** is used to decide the **order of filling of orbitals**.

- n= principal quantum number
- l= azimuthal (orbital) quantum number

Rule:

1. The orbital with the **lower value of (n + l)** is filled first.

2. If two orbitals have the same (n + l) value, the one with the **lower n** fills first.

9. Which of the following orbits will be filled first?

a) 2p or 3s

b) 3d or 4s

Ans: **(a) 2p vs 3s**

- $2\text{p} \rightarrow n = 2, l = 1 \Rightarrow n + l = 3$
- $3\text{s} \rightarrow n = 3, l = 0 \Rightarrow n + l = 3$

Same value → orbital with **lower n fills first**

Answer: 2p fills first

(b) 3d vs 4s

- $3\text{d} \rightarrow n = 3, l = 2 \Rightarrow n + l = 5$
- $4\text{s} \rightarrow n = 4, l = 0 \Rightarrow n + l = 4$

Lower value fills first

Answer: 4s fills first

10. The electronic configuration of Cr is (Ar) $3d^5 4s^1$ not $3d^4 4s^2$

Ans: The actual electronic configuration of chromium (Cr) is:

(Ar) $3d^5$ $4s^1$ (not $3d^4$ $4s^2$)

Reason: Extra Stability of Half-Filled Orbitals

- A **half-filled d-subshell ($3d^5$)** is especially stable.
- This stability comes from:
  - **Symmetrical distribution** of electrons
  - **Exchange energy** (more possible electron exchanges with parallel spins)

Instead of:

- $3d^4$ $4s^2$

One electron from 4s shifts to 3d, giving:

- **$3d^5$ $4s^1$**

This results in a **more stable arrangement**.

4

Chemical Bonding

1. Core Concepts

Octet Rule: Atoms combine to achieve a stable configuration of 8 electrons in their valence shell.

Ionic Bond: Formed by the complete transfer of electrons from a metal to a non-metal **(e.g., NaCl)**

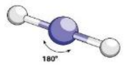
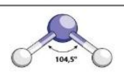
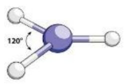
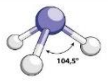
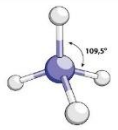
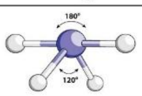
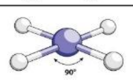
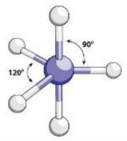
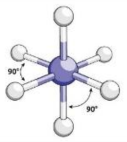
Covalent Bond: Formed by the mutual sharing of electrons between non-metals.

Hybridization: The mixing of atomic orbitals of slightly different energies to form new equivalent orbitals **(e.g., sp, sp^2, sp^3)**

2. VSEPR Theory

Predicts the shape of molecules based on electron pair repulsion around the central atom.

Bond Pair - Bond Pair < Bond Pair - Lone Pair < Lone Pair - Lone Pair repulsion order.

Molecular Formula	VSEPR Notation	Molecular Geometry	Geometric Shape	Example
AX_2	AX_2		Linear	CO_2
	AX_2E_2		Bent(V-shaped)	H_2O, OF_2
AX_3	AX_3		Trigonal Planar	BH_3, SO_3
	AX_3E		Trigonal Pyramidal	NH_3, PCl_3
AX_4	AX_4		Tetrahedral	CH_4
	AX_4E		Seesaw	SF_4
	AX_4E_2		Square Planar	XeF_4
AX_5	AX_5		Trigonal Bipyramidal	PCl_5
AX_6	AX_6		Octahedral	SF_6

3. Valence Bond Theory (VBT)

A. Key Concepts

Overlapping: When two atoms' half-filled atomic orbitals combine, a covalent bond is formed.

Electron Spin: The spins of the two overlapping electrons must always be opposite.

Strength of Bond: The greater the overlap, the stronger the bond will be.

B. Types of Overlapping (Bonds)

Sigma (σ) Bond:

Formed when orbitals overlap head-on (**directly**) or axially.

This is a strong bond.

Examples: s-s, s-p, and **p-p** (**axial**) overlap.

Pi (π) Bond:

Formed when **p**-orbitals overlap sideways (**laterally**).

Weaker than a **σ** bond.

Always formed after a **σ** bond has already been established.

Feature	Valence Bond Theory (VBT)	Molecular Orbital Theory (MOT)
Electron Location	Electrons remain localized around their own atom.	Electrons are delocalized across the entire molecule.
Bonding	Formed by overlap of half-filled orbitals.	Formed by merging atomic orbitals to create molecular orbitals (MOs).
Stability	Cannot explain the paramagnetism of O_2.	Successfully explains that O_2 is paramagnetic.

4. Molecular Orbital Theory (MOT) - Detailed

MOT explains that atomic orbitals combine to form **Molecular Orbitals (MO)** which belong to the entire molecule.

A. Types of Molecular Orbitals

1. **Bonding MO (BMO):** Lower energy, higher stability **(σ, π).**
2. **Anti-bonding MO (ABMO):** Higher energy, lower stability **(σ*, π*)**

B. Electronic Configurations (Energy Order)

The sequence of filling electrons depends on the total electron count:

Case 1: For Molecules with ≤ 14 Electrons (e.g., Li_2, B_2, C_2, N_2)

Due to **s-p** mixing, **σ2pz** has higher energy than **$\pi 2p_x$** and **$\pi 2p_y$**

Order: $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

Case 2: For Molecules with > 14 Electrons (e.g., O_2, F_2)

$\sigma 2p_z$ is lower in energy than the **π** orbitals.

Order: $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

C. Bond Order

It is half the difference between electrons in Bonding **(Nb)** and Anti-bonding **(Na)** orbitals.

Formula: $\text{Bond Order} = \frac{1}{2}(N_b - N_a)$

Stability: If Bond Order is **> 0,** the molecule is stable.

Magnetic Nature: Unpaired electrons = **Paramagnetic**; All paired = **Diamagnetic**.

Top 10 Important Questions

Q1. Calculate the Bond Order of N_2 (14 electrons).

Ans: $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 (\pi 2p_x^2 = \pi 2p_y^2) \sigma 2p_z^2$

$N_b = 10, N_a = 4 \rightarrow \text{B.O.} = \frac{10-4}{2} = \mathbf{3}.$

Q2. Why is O_2 Paramagnetic?

Ans: According to **MOT, O_2** has two unpaired electrons in the anti-bonding **$\pi^* 2p_x$** and **$\pi^* 2p_y$** orbitals.

Q3. Explain the geometry of H_2O molecule.

Ans: Oxygen has **2** bond pairs and **2** lone pairs. According to **VSEPR**, the shape is **Bent/V-shaped** with an angle of **104.5°**

Q4. What is the hybridization of Carbon in Ethene (C_2H_4)?

Ans: Each Carbon is **sp^2 hybridized**, forming a planar structure with a **120°** angle.

Q5. Which one the pairs is more covalent and why?

a) AgCl, AgI
b) LiCl, KCl

Ans: **(a) Which one is more covalent and why?**

(i) AgCl vs AgI → AgI is more covalent

- According to **Fajan's Rule**, larger anions are more polarizable.
- I^- is larger than Cl^- → more distortion of electron cloud → more covalent character.
- Hence, **AgI > AgCl (more covalent)**

(ii) LiCl vs KCl → LiCl is more covalent

- Smaller cation = higher polarizing power.
- Li^+ is smaller than K^+ → distorts Cl^- more.
- Hence, **LiCl > KCl (more covalent)**

Q6. Predict the molecular geometrics of AX_4 Type molecular having :

a) 4 bond pairs and no lone pairs
b) 2 bond pairs and 2 lone pairs

Ans: **(b) Molecular geometry of AX_4 type molecules**

Using **VSEPR Theory**:

(i) 4 bond pairs, 0 lone pairs → Tetrahedral

- Example: CH_4
- Bond angle ≈ 109.5°

Q7. Be_2 molecular geometrics of AX_4 types molecules having:

Ans: **(c) Be_2 molecule does not exist (MO Theory)**

- Electronic configuration of Be = $1s^2\ 2s^2$
- Total electrons in Be_2 = 8

Filling Molecular Orbitals:

- $\sigma 1s^2$, $\sigma 1s^{*2}$, $\sigma 2s^2$, $\sigma 2s^{*2}$

Bond Order = (Nb − Na)/2 = (4 − 4)/2 = 0

- Bond order = 0 → no stable bond forms

Therefore, Be2 does not exist

Q8. Complete the following choosing from the given option:

(one, two, zero, 4, 2, 3)

a) **Be2 is not formed because its bond order is ....**
b) **The number of molecular orbitals formed by mixing two atomic orbitals from each of two atoms is ....**

Ans: a) Zero, b) two

Q9. Write true (T) for correct statement and False (F) for incorrect statement.

a) **Bond order of O^{-}_{2} molecules is 2**
b) **Resonance hybrid of carbonate ion has 3 contributing canonical structure**

Ans: (a) For O_2^-, the bond order is **1.5**, not 2 → **False (F)**

(b) CO_3^{2-} has **3 contributing resonance (canonical) structures** → **True (T)**

Q10. State the posulates of valence shell electron pair repulsion theory. What is the order order of repulsive forces between different type of electron pair? What is the expected geometry of molecules AX_4, AX_5 and AX_6 ? give one example of each

Ans: Postulates of VSEPR theory

1. Electron pairs (bond pairs + lone pairs) around a central atom repel each other.
2. These electron pairs arrange themselves to minimize repulsion and achieve maximum separation.
3. The shape of a molecule depends on the number of electron pairs around the central atom.
4. Lone pairs occupy more space than bond pairs.
5. Multiple bonds behave as a single electron pair region but repel more than single bonds.

Order of repulsive forces

Lone pair – Lone pair > Lone pair – Bond pair > Bond pair – Bond pair

Geometries of molecules

- **AX_4 → Tetrahedral**
  Example: CH_4
- **AX_5 → Trigonal bipyramidal**
  Example: PCl_5
- **AX_6 → Octahedral**
  Example: SF_6

7

Solutions

1. Introduction & Concentration Terms

Solution: A homogeneous mixture of two or more substances. The substance in larger quantity is the **solvent**, and the smaller is the **solute**.

Molarity (M): Moles of solute per litre of solution. $M = \frac{n_{solute}}{V_{solution(L)}}$.

Molality (m): Moles of solute per kg of solvent. $m = \frac{n_{solute}}{W_{solvent(kg)}}$.

Mole Fraction (x): Ratio of moles of one component to the total moles. $x_A = \frac{n_A}{n_A+n_B}$.

2. Solubility & Henry's Law

Henry's Law: The solubility (**mass**) of a gas in a liquid is directly proportional to its partial pressure. **$m = k \cdot P$**

3. Vapour Pressure & Raoult's Law

Vapour Pressure: The pressure exerted by vapours in equilibrium with the liquid at a constant temperature.

Raoult's Law: For a solution of volatile liquids, the partial vapour pressure of each component is proportional to its mole fraction. **$P_A = P°_A \cdot x_A$**

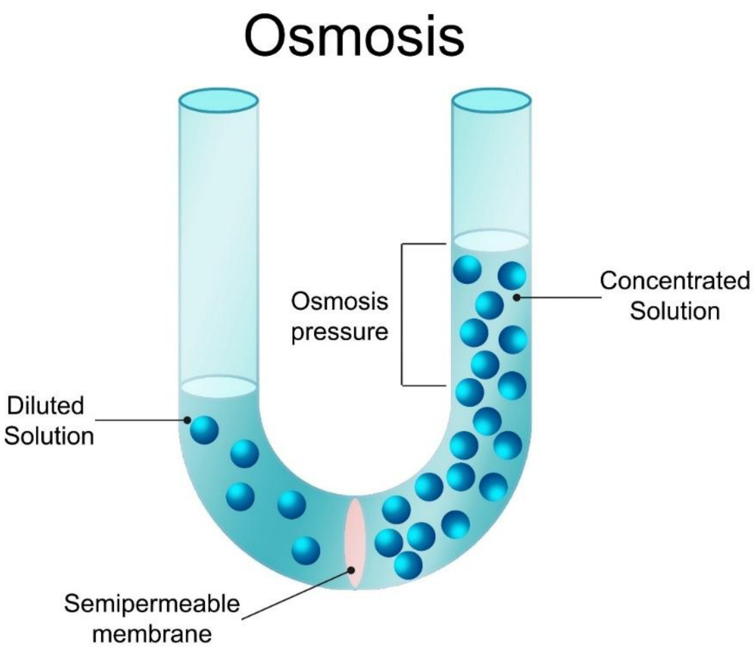
Ideal Solution: Follows Raoult's law at all concentrations (e.g., Benzene + Toluene).

Non-Ideal Solution: Shows positive or negative deviations from Raoult's Law.

4. Colligative Properties

These depend only on the **number of solute particles**, not their nature.

1. **Relative Lowering of Vapour Pressure:** $\frac{P^\circ - P}{P^\circ} = x_{solute}$.
2. **Elevation of Boiling Point (ΔTb):** Boiling point increases when a non-volatile solute is added. **$\Delta T_b = K_b \cdot m$**
3. **Depression of Freezing Point (ΔT_f):** Freezing point decreases. **$\Delta T_f = K_f \cdot m$**
4. **Osmotic Pressure (α):** The pressure required to stop osmosis. **$\pi = CRT$**



5. Van't Hoff Factor

Used for solutes that undergo association or dissociation.

$$i = \frac{\text{Observed Colligative Property}}{\text{Calculated Colligative Property}}.$$

10 expected Questions

Q1. Define Molality and explain why it is preferred over Molarity.

Ans. Molality (m):

It is the number of moles of solute dissolved in 1 kg of solvent.

Formula:

$$m = \frac{\text{moles of solute}}{\text{mass of solvent in kg}}$$

Molality is independent of temperature because it depends on mass, not volume.

Molarity changes with temperature (due to expansion/contraction of volume), but molality remains constant.

Q2. State Henry's Law and its two limitations.

Ans. Henry's Law:

At constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of that gas above the liquid.

Two limitations:

1. It is valid only for low pressure conditions.
2. It does not apply when the gas reacts chemically with the solvent.

Q3. Calculate the mole fraction of ethylene glycol in a 20% by mass solution.

Ans. 20% by mass → 20 g ethylene glycol + 80 g water

Moles of ethylene glycol = $\frac{20}{62} \approx 0.323$

Moles of water = $\frac{80}{18} \approx 4.44$

Mole fraction of ethylene glycol:

$$X = \frac{0.323}{0.323 + 4.44} \approx 0.068$$

= Mole fraction ≈ 0.068

Q4. What is the difference between Ideal and Non-ideal solutions?

Ans. Ideal Solution:

A solution that obeys Raoult's Law at all concentrations and has no heat or volume change on mixing.

Non-ideal Solution:

A solution that does not obey Raoult's Law and shows change in heat and/or volume on mixing.

Q5. Why does the boiling point of water increase when salt is added?

Ans. Adding salt (a non-volatile solute) lowers the vapour pressure of water.

As a result, a higher temperature is required to reach boiling point, so the boiling point increases.

Q6. a) State Henry's law. Give its expression.

b) Mention any two condition under which gases obey Henry's law

Ans. (i) Henry's Law:

At constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.

Expression:

$$p = k_H x$$

(where p= partial pressure, x= mole fraction, k_H= Henry's constant)

(ii) Two conditions:

1. **Low pressure**
2. **No chemical reaction** between gas and solvent

Q7. The vapour pressure of methanol and ethanol are 88.7 mm and 44.5 mm of Hg respectively. A solution is prepared by mixing 60 g ethanol and 40 g of methanol. Assuming the solution to be ideal, calculate the vapour pressure of the solution.

Ans. mm Hg

$P_{\text{ethanol}} = 44.5\text{mm Hg}$

Mass: ethanol = 60 g, methanol = 40 g

Moles:

Ethanol = $\frac{60}{46} \approx 1.304$

Methanol = $\frac{40}{32} = 1.25$

Total moles = $1.304 + 1.25 = 2.554$

Mole fractions:

$$X_{\text{ethanol}} = \frac{1.304}{2.554} \approx 0.51$$
$$X_{\text{methanol}} = \frac{1.25}{2.554} \approx 0.49$$

Using Raoult's Law:

$$P_{\text{solution}} = (0.51 \times 44.5) + (0.49 \times 88.7)$$
$$P = 22.7 + 43.5 \approx 66.2 \text{ mm Hg}$$

= 66.5 mm Hg

Q8. Calculate the normality of a solution of NaOH if 0.4 g of NaOH is dissolved in 100 ml of the solution.

Ans. Mass of NaOH = 0.4 g
Volume = 100 mL = 0.1 L

Equivalent weight of NaOH = 40

Normality (N):

$$N = \frac{\text{mass}}{\text{eq. wt.} \times \text{volume in L}} = \frac{0.4}{40 \times 0.1}$$
$$N = \frac{0.4}{4} = 0.1$$

= 0.1 N

Q9. What is meant by Osmosis and reverse osmosis? Write one application of reverse Osmosis.

Ans. Osmosis:
The movement of solvent from a dilute solution to a concentrated solution through a semipermeable membrane.

Reverse Osmosis:
The flow of solvent is reversed (from concentrated to dilute solution) by applying external pressure greater than osmotic pressure.

One application:
Used in water purification (RO systems) to remove impurities from drinking water.

Q10. A solution containing 12.5 g of a non-electrolyte substance in 175 g of water gave boiling point elevation of 0.70 k calculate the molar mass of the substance.

Ans. $W = 175\, g = 0.175\, kg$
$\Delta T_b = 0.70\, K,\ K_b = 0.52\, K\, kg/mol$(for water)

Formula:

$$\Delta T_b = K_b \times \frac{w}{M \times W}$$
$$0.70 = 0.52 \times \frac{12.5}{M \times 0.175}$$
$$M = \frac{0.52 \times 12.5}{0.70 \times 0.175}$$
$$M = \frac{6.5}{0.1225} \approx 53\, g/mol$$

= 53 g/mol

9

Chemical Thermodynamics

1. Fundamental Terms

System and Surroundings: The part of the universe under study is the **system**, and the rest is the **surroundings**.

Types of Systems:

Open: Can exchange both energy and matter.

Closed: Can exchange energy but not matter.

Isolated: Can exchange neither energy nor matter.

State Functions: Properties that depend only on the current state, not the path taken **(e.g., P, V, T, U, H)**

2. First Law of Thermodynamics

Energy Conservation: Energy can neither be created nor destroyed, only transformed.

Mathematical Expression: $\Delta U = q + w$

ΔU = Change in Internal Energy

q = Heat added to the system

w = Work done on the system.

3. Enthalpy (H)

It is the total heat content of a system at constant pressure.

Relation with Internal Energy: $\Delta H = \Delta U + \Delta n_g RT$

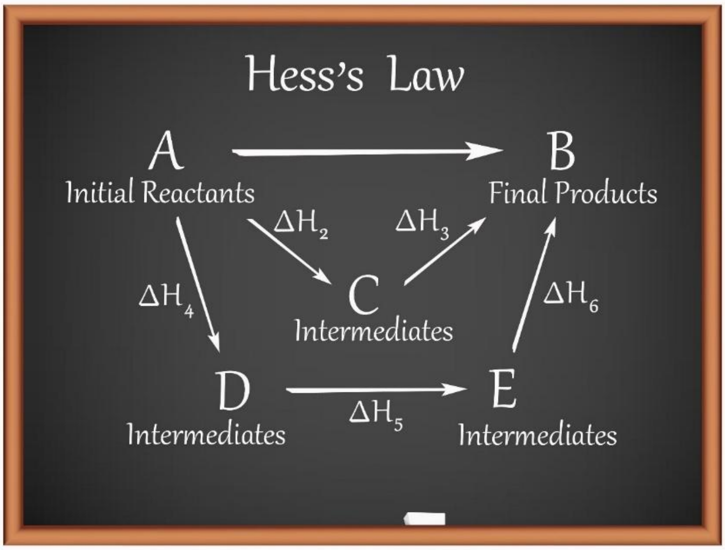
Δn_g = Moles of gaseous products - Moles of gaseous reactants.

4. Thermochemistry

Exothermic: Heat is evolved (ΔH is negative).

Endothermic: Heat is absorbed (ΔH is positive).

Hess's Law of Constant Heat Summation: The total enthalpy change for a reaction is the same whether it occurs in one step or several steps.



5. Standard Enthalpies

Enthalpy of Formation ($\Delta_f H°$): Change when **1** mole of a substance is formed from its elements in their standard states.

Enthalpy of Neutralization: Heat change when **1 g**-equivalent of an acid is neutralized by a base.

10 Expected Questions

Q1. Define Open, Closed, and Isolated systems with one example each.

Ans. Open System:

A system that can exchange both matter and energy with surroundings.

Example: Boiling water in an open vessel.

Closed System:

A system that can exchange only energy, not matter with surroundings.

Example: Water in a closed bottle.

Isolated System:

A system that exchanges neither matter nor energy with surroundings.

Example: Thermos flask (ideal case).

Q2. State the First Law of Thermodynamics and write its mathematical form.

Ans. First Law of Thermodynamics:

Energy can neither be created nor destroyed; it only changes from one form to another.

Mathematical form:

$$\Delta U = Q - W$$

Q3. Calculate Δ H for the reaction: $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ using standard enthalpies of formation.

Ans. Formula:

$$\Delta H = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$$

Values (kJ/mol):

C_2H_4 = +52.3, O_2 = 0, CO_2 = –393.5, $H_2O(l)$ = –285.8

$$\Delta H = [2(-393.5) + 2(-285.8)] - [52.3 + 0]$$
$$= (-787 - 571.6) - 52.3 = -1410.9 \text{ kJ}$$

= -1410.9 kJ

Q4. Distinguish between State Functions and Path Functions.

Ans. State Functions:

Depend only on the initial and final state, not on the path taken.

Examples: Internal energy (U), Enthalpy (H), Pressure, Temperature.

Path Functions:

Depend on the path taken during the process.

Examples: Heat (Q), Work (W).

Q5. What is the sign of Δ H for an endothermic reaction? Give an example.

Ans. Sign of ΔH:

For an endothermic reaction, ΔH is positive (+).

Example:

Melting of ice:

H_2O (s) → H_2O (l)

Q6. Define isothermal process. Distinguish between the Adiabatic and isothermal processes.

Ans. **Isothermal process:**
An isothermal process is a thermodynamic process in which the temperature of the system remains constant throughout ($\Delta T = 0$).

Difference between isothermal and adiabatic processes:

- **Heat exchange:**
  Isothermal → Heat is exchanged with surroundings
  Adiabatic → No heat exchange ($Q = 0$)
- **Temperature:**
  Isothermal → Constant temperature
  Adiabatic → Temperature changes
- **Energy change:**
  Isothermal → Internal energy remains constant
  Adiabatic → Internal energy changes
- **Example:**
  Isothermal → Slow expansion of gas
  Adiabatic → Rapid compression/expansion of gas

Q7. Which of the following will increase the internal energy of a system?

a) **Heat given to the system**
b) **Work done by the system**

Ans. **(a) Heat given to the system**
(b) Work done by the system → decreases internal energy

Q8. Calculate the enthalpy change in the reaction.

$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(l)$ at 298 K. Given that enthalpy of formation for $NH_3(g)$ and H_2O (l) are -46.0 kJ mol^{-1} and 286.0 kJ mol^{-1} respectively. Also mention the standard enthalpy of formation of an element.

Ans. **Use:**

$$\Delta H^\circ_{\text{reaction}} = \sum \nu \Delta H^\circ_f(\text{products}) - \sum \nu \Delta H^\circ_f(\text{reactants})$$

Given reaction:

$$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(l)$$

Data:

$\Delta H^\circ f(NH_3) = -46.01$ kJ mol^{-1}

$\Delta H^\circ f(H_2O(l)) = -286.0$ kJ mol^{-1}

$\Delta H^\circ f(N_2) = 0$ (element)

$\Delta H°f(O_2)$ = 0 (element)

Step calculation:

Products:

$2(0) + 6(-286.0) = -1716.0$ kJ

Reactants:

$4(-46.01) + 3(0) = -184.04$ kJ
$\Delta H° = -1716.0 - (-184.04) = -1531.96$ kJ

= $\Delta H° = -1532$ kJ

Q9. The enthalpies of formation of OH(g), H(g) and O(g) are 42, 218 and 248 kJ mol^{-1} Calculate the bond enthalpy of O-H.

Ans. Given enthalpies of formation (ΔHf):

ΔHf[OH(g)] = 42 kJ mol^{-1}

ΔHf[H(g)] = 218 kJ mol^{-1}

ΔHf[O(g)] = 248 kJ mol^{-1}

The O–H bond enthalpy is the energy required to break:

OH(g) → O(g) + H(g)

Use the relation:

$\Delta H = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$
$\Delta H = [\Delta H_f(O) + \Delta H_f(H)] - \Delta H_f(OH)$

Substitute values:

$\Delta H = (248 + 218) - 42$
$\Delta H = 466 - 42$
$\Delta H = 424$ kJ mol^{-1}

Bond enthalpy of O–H = 424 kJ mol^{-1}

Q10. State Lavoiser – Law of thermochemistry with suitable example.

Ans. Lavoisier–Laplace Law of Thermochemistry

Statement:
The heat change (enthalpy change) of a chemical reaction is equal in magnitude but opposite in sign to the heat change of the reverse reaction.

Formation of water:

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \Delta H = -286 \text{ kJ mol}^{-1}$

This means 286 kJ heat is released.

Reverse reaction:

$H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g) \Delta H = +286 \text{ kJ mol}^{-1}$

Here 286 kJ heat is absorbed.

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Ionic Equilibrium

1. Electrolytes

Strong Electrolytes: Substances that ionize almost completely in aqueous solution **(e.g., HCl, NaOH, NaCl).**

Weak Electrolytes: Substances that ionize only partially in aqueous solution **(e.g., CH_3COOH, NH_4OH).**

Degree of Ionization (α): The fraction of the total number of molecules of an electrolyte that ionize in solution.

2. Acid-Base Concepts

Arrhenius Concept: Acids give **H^+** ions; Bases give **OH^-** ions in water.

Bronsted-Lowry Concept: Acids are proton **(H^+)** donors; Bases are proton acceptors.

Conjugate Acid-Base Pair: A pair of substances that differ only by one proton.

Lewis Concept: Acids are electron-pair acceptors; Bases are electron-pair donors.

3. Ionization of Water and pH Scale

Ionic Product of Water (K_w): At 298 K, $K_w = [H3O^+] [OH^-] = 1.0 \times 10^{-14}$

pH Scale: $pH = -\log [H_3O^+]$

Neutral: pH = 7; Acidic: pH < 7; Basic: pH > 7

4. Ostwald's Dilution Law

For a weak electrolyte, the degree of ionization **(α)** is inversely proportional to the square root of its molar concentration **(c):**

$$\alpha = \sqrt{\frac{K}{c}}.$$

5. Buffer Solutions and Salt Hydrolysis

Buffer Solution: A solution that resists change in **pH** on the addition of small amounts of acid or base.

Common Ion Effect: The suppression of the degree of ionization of a weak electrolyte by the addition of a strong electrolyte having a common ion.

Solubility Product (K_{sp}): The product of the molar concentrations of ions in a saturated solution of a sparingly soluble salt

Top 10 Expected Questions

Q1. Differentiate between strong and weak electrolytes with examples.

Ans.

Basis	Strong Electrolytes	Weak Electrolytes
Definition	Electrolytes that completely ionize in aqueous solution.	Electrolytes that partially ionize in aqueous solution.
Ionization	Nearly 100% dissociation into ions.	Incomplete dissociation into ions.
Electrical Conductivity	High conductivity because many ions are present.	Low conductivity because fewer ions are present.
Nature of equilibrium	Ionization is almost irreversible.	Ionization is reversible and establishes equilibrium.
Examples	HCl, HNO_3, NaCl, KOH	CH_3COOH, NH_4OH, H_2CO_3

Q2. Explain the Bronsted-Lowry concept of acids and bases using the example of NH_3 and H_2O

Ans. Bronsted–Lowry Concept of Acids and Bases

Definition:

Bronsted–Lowry Acid: A substance that donates a proton (H^+).

Bronsted–Lowry Base: A substance that accepts a proton (H^+).

Example: Reaction between NH_3 and H_2O

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

Explanation:

NH_3 (Ammonia) accepts a proton (H^+) from water → acts as a base.

H_2O (Water) donates a proton (H^+) to ammonia → acts as an acid.

Products formed:

NH_4^+ (Ammonium ion) → conjugate acid of NH_3

OH^- (Hydroxide ion) → conjugate base of H_2O

Q3. Define a Conjugate Acid-Base pair. Give the conjugate base for HCO_3^- and conjugate acid for NH_3

Ans. Conjugate Acid–Base Pair:

A pair of substances that differ by one proton (H^+). The acid loses H^+ to form its conjugate base, and the base gains H^+ to form its conjugate acid.

Conjugate base of HCO_3^-: CO_3^{2-}

Conjugate acid of NH_3: NH_4^+.

Q4. Derive the expression for Ostwald's Dilution Law for a weak acid.

Ans. For a weak acid HA:

$$HA \rightleftharpoons H^+ + A^-$$

Let initial concentration = C, degree of dissociation = α.

At equilibrium:

$[HA] = C(1-\alpha)$, $[H^+] = C\alpha$, $[A^-] = C\alpha$

Acid dissociation constant:

$$K_a = \frac{[H^+][A^-]}{[HA]}$$
$$K_a = \frac{(C\alpha)(C\alpha)}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha}$$

For weak acids α is very small, so $1-\alpha \approx 1$

$$K_a = C\alpha^2$$
$$\alpha = \sqrt{\frac{K_a}{C}}$$

This is Ostwald's Dilution Law.

Q5. Calculate the pH of a 1.0 x 10^{-4} M solution of NaOH.

Ans. NaOH is a strong base, so it fully dissociates:

$$[OH^-] = 1.0 \times 10^{-4}\ M$$
$$pOH = -\log(1.0 \times 10^{-4}) = 4$$
$$pH = 14 - pOH = 14 - 4 = 10$$

Q6. What is the pH of a 0.01 M aqueous solution of HCL?

Ans. HCl is a strong acid, so it completely dissociates.

$[H^+] = 0.01 = 1 \times 10^{-2}$
$pH = -\log[H^+]$
$pH = -\log(10^{-2}) = 2$

Answer: pH = 2

Q7. Calculate the pH of ammonium hydroxide – ammonium chloride buffer solution that is 0.1 M in ammonium hydroxide and 0.01 M in ammonium chloride. (pK_b of NH_4OH = 9.25)

Ans. For a basic buffer (NH_4OH / NH_4Cl):

$$pOH = pK_b + \log\frac{[\text{salt}]}{[\text{base}]}$$

Given:

$pK_b = 9.25, [NH_4Cl] = 0.01\,M, [NH_4OH] = 0.1\,M$
$pOH = 9.25 + \log\frac{0.01}{0.1}$
$pOH = 9.25 + \log(0.1)$
$\log(0.1) = -1$
$pOH = 9.25 - 1 = 8.25$
$pH = 14 - pOH = 14 - 8.25 = 5.75$

Answer: pH = 5.75

Q8. Calculate the solubility product for Ag_2SO_4 if [$SO4^{2-}$] = 2.5 ' 10^{-2} M

Ans. Dissociation:

$Ag_2SO_4 \rightleftharpoons 2Ag^+ + SO_4^{2-}$

Given:

$[SO_4^{2-}] = 2.5 \times 10^{-2} M$
$[Ag^+] = 2 \times 2.5 \times 10^{-2} = 5.0 \times 10^{-2} M$
$K_{sp} = [Ag^+]^2[SO_4^{2-}]$
Ksp = 6.25×10^{-5}.

Q9. Explain the buffer action of sodium acetate and acetic acid buffer solution.

Ans. Buffer Action of Acetic Acid – Sodium Acetate

A mixture of acetic acid (CH_3COOH) and sodium acetate (CH_3COONa) forms an acidic buffer solution that resists changes in pH.

Action:

1. When acid (H^+) is added:
Acetate ions remove the added H^+.

$CH_3COO^- + H^+ \rightarrow CH_3COOH$

2. When base (OH^-) is added:
Acetic acid neutralizes the added OH^-.

$$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$$

Q10. Derive the expression for the degree of dissociation of weak acid. What is the effect of common ion on the degree of dissociation.

Ans. Degree of Dissociation of Weak Acid

For weak acid **HA**:

$$HA \rightleftharpoons H^+ + A^-$$
$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}$$

For weak acids **α is very small**, so **1−α ≈ 1**

$$K_a = C\alpha^2$$
$$\alpha = \sqrt{\frac{K_a}{C}}$$

Effect of Common Ion:
Addition of a **common ion decreases the degree of dissociation** of a weak acid.

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Electrochemistry

1. Oxidation and Reduction

Oxidation: Loss of electrons by an atom or ion.

Reduction: Gain of electrons by an atom or ion.

Redox Reaction: **A** reaction where both oxidation and reduction occur simultaneously.

Oxidation Number (ON): The charge an atom appears to have when all other atoms are removed from it as ions.

2. Electrolytic Conduction

Electrolytes: Substances that conduct electricity in their aqueous or molten state by the movement of ions.

Conductance (G): The ease with which current flows through a conductor **(G = 1/R).** Unit: Siemens **(S) or Ω -1.**

Molar Conductivity (λ_m): The conducting power of all the ions produced by dissolving one mole of an electrolyte in solution.

Kohlrausch's Law: At infinite dilution, the molar conductivity of an electrolyte is the sum of the individual contributions of its anions and cations.

3. Electrolysis and Faraday's Laws

Electrolysis: The process of decomposition of an electrolyte by passing electricity.

Faraday's First Law: The mass of a substance deposited or liberated at an electrode is directly proportional to the quantity of electricity passed **(m = ZIt).**

Faraday's Second Law: When the same quantity of electricity is passed through different electrolytes, the masses of substances liberated are proportional to their chemical equivalent weights.

4. Electrochemical Cells

Electrolytic Cell: Converts electrical energy into chemical energy (non-spontaneous reactions).

Galvanic (Voltaic) Cell: Converts chemical energy into electrical energy (spontaneous reactions).

Standard Hydrogen Electrode (SHE): Used as a reference electrode with a potential assigned as **0.00 V.**

5. Nernst Equation

It relates the electrode potential to the concentration of ions:

$$E = E^\circ - \frac{0.0591}{n} \log \frac{[Product]}{[Reactant]}$$

Electrochemical Series: Arrangement of electrodes in increasing order of their standard reduction potentials.

6. Fundamental Concepts

Oxidation & Reduction: Oxidation is the loss of electrons **($Zn \rightarrow Zn^{2+} + 2e^-$);** Reduction is the gain of electrons **($Cu^{2+} + 2e^- \rightarrow Cu$).**

Oxidation Number: The formal charge an atom carries in a compound.

7. Gibbs Free Energy and EMF (Δ G) - New

The electrical work done by a Galvanic cell is equal to the decrease in its Gibbs free energy.

Formula: **$\Delta G = -nFE_{cell}$**

n = number of moles of electrons transferred.

F = Faraday's constant **(96500 C/mol)**

E_{cell} = EMF of the cell.

Spontaneity: For a cell to work spontaneously, **Δ G** must be **negative**, which means **E_{cell}** must be **positive**.

8. Types of Commercial Cells (Batteries) - New

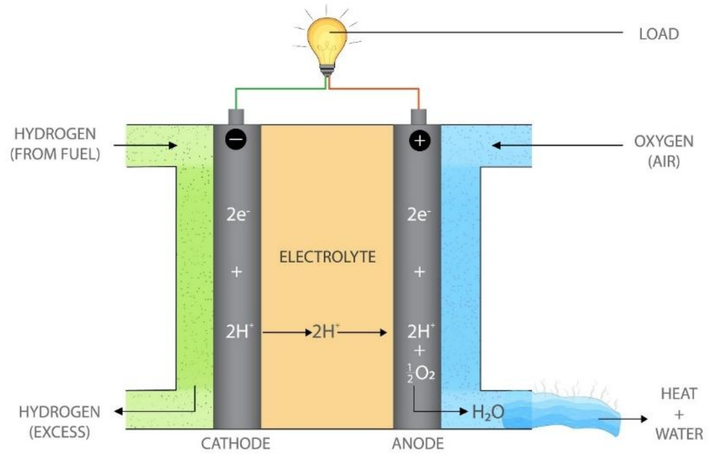
Batteries are categorized into two main types based on their rechargeability:

Primary Cells: The reaction occurs only once; they cannot be recharged (e.g., **Dry Cell** or Leclanche cell used in clocks).

Secondary Cells: These can be recharged by passing current in the opposite direction (e.g., **Lead-Acid Storage Battery** used in cars, and Nickel-Cadmium cells).

Fuel Cells: Cells that convert the energy of combustion of fuels like Hydrogen directly into electrical energy. They are pollution-free and highly efficient (e.g., **H_2-O_2** Fuel Cell).

FUEL CELL



9. Corrosion - New

Definition: The slow destruction of metals due to their reaction with environmental gases and moisture (e.g., **Rusting of Iron**).

Mechanism: It is an electrochemical phenomenon. At the **Anode**, Iron is oxidized **($Fe \rightarrow Fe^{2+} + 2e^-$).** At the **Cathode**, Oxygen is reduced in the presence of **H^+** ions.

Prevention: Coating with paint, bisphenols, or **Galvanization** (coating iron with a layer of Zinc).

Top 10 Expected Questions

Q1. Define Oxidation Number. Calculate the ON of S in H_2SO_4

Ans. Oxidation Number (ON):

The oxidation number is the apparent charge assigned to an atom in a molecule or ion according to certain rules.

Calculation for H_2SO_4:

Let oxidation number of S = x

H = +1, O = −2

$$2(+1) + x + 4(-2) = 0$$
$$2 + x - 8 = 0$$
$$x - 6 = 0$$
$$x = +6$$

Oxidation number of S in H_2SO_4 = +6.

Q2. State Faraday's First and Second laws of electrolysis.

Ans. Faraday's First Law of Electrolysis:

The mass of a substance deposited or liberated at an electrode is directly proportional to the quantity of electricity passed through the electrolyte.

$m \propto Q$

Faraday's Second Law of Electrolysis:

When the same quantity of electricity passes through different electrolytes, the masses of substances deposited are proportional to their equivalent weights.

Q3. Explain the effect of dilution on Molar Conductivity for strong and weak electrolytes.

Ans. Effect of Dilution on Molar Conductivity (Λm):

Strong Electrolytes:

Molar conductivity increases slightly with dilution because inter-ionic attraction decreases and ions move more freely.

Weak Electrolytes:

Molar conductivity increases sharply with dilution because the degree of ionization increases, producing more ions.

Q4. Differentiate between an Electrolytic cell and a Galvanic cell.

Ans.

Basis	Electrolytic Cell	Galvanic Cell
Energy conversion	Electrical energy → Chemical energy	Chemical energy → Electrical energy
Nature of reaction	Non-spontaneous reaction	Spontaneous reaction
Power source	Requires external power supply	No external power needed
Example	Electrolysis of water	Daniell cell

Q5. State Kohlrausch's Law and give one of its applications.

Ans. Kohlrausch's Law:
At infinite dilution, each ion contributes independently to the molar conductivity of an electrolyte.

$\Lambda_m^\circ = \lambda_+^\circ + \lambda_-^\circ$

Application:
It is used to calculate the molar conductivity of weak electrolytes at infinite dilution

Q6. Predict whether the following reaction is feasible or not?

Cu^{2+} (aq) + 2Ag(s) → Cu(s) + $2Ag^+$ (aq)

$E^\circ_{Ag^+/Ag}$ = 0.80 V and $E^\circ_{Cu^{2+}/Cu}$ = 0.34 V

Ans. Standard reduction potentials:

- $E^\circ_{Ag^+/Ag} = +0.80\, V$
- $E^\circ_{Cu^{2+}/Cu} = +0.34\, V$

In the reaction:

- **Cu^{2+}** → **Cu** (reduction) → $E^\circ = +0.34\, V$
- **Ag** → **Ag^+** (oxidation) → $E^\circ = -0.80\, V$

Cell potential:

$$E^\circ_{cell} = E^\circ_{reduction} + E^\circ_{oxidation}$$
$$E^\circ_{cell} = 0.34 + (-0.80) = -0.46\, V$$

Since $E^\circ_{cell} < 0$, the reaction is **not feasible (non-spontaneous)**.

= $E^\circ_{cell} = -0.46\, V$, so the reaction is **not feasible**.

Q7. Calculate the reduction potential for the following half cell at 298 K

Ag^+ (0.1M) + e^- → Ag (s)

E^0 = 0.80V

Ans. $[Ag^+] = 0.1\, M,\, T = 298K$

Using **Nernst equation**:

$$E = E^\circ - \frac{0.0591}{n} \log \frac{1}{[Ag^+]}$$

Here $n = 1$

$$E = 0.80 - 0.0591 \log \frac{1}{0.1}$$

$\log(10) = 1$
$E = 0.80 - 0.0591$
$E = 0.7409\,V$

Reduction potential ≈ 0.74 V.

Q8. Write the cell reaction, Nernst equation and calculate the EMF of following cell at 298k

Fe(s) | Fe^{2+}(0.001 M) || H^+(1 M) | $H_{2(g)}$ (1 bar) | Pt(s)

[Given : $E°_{Fe2+/Fe}$ = -0.44 V]

Ans. **Half reactions:**
Anode: $Fe \rightarrow Fe^{2+} + 2e^-$
Cathode: $2H^+ + 2e^- \rightarrow H_2$

Cell reaction:

$Fe + 2H^+ \rightarrow Fe^{2+} + H_2$

Standard EMF:

$E^\circ_{cell} = 0 - (-0.44) = 0.44\,V$

Nernst equation:

$$E = E^\circ - \frac{0.0591}{2}\log\frac{[Fe^{2+}]}{[H^+]^2}$$
$$E = 0.44 - \frac{0.0591}{2}\log(0.001)$$
$$E \approx 0.53\,V$$

Q9. How is a bridge formed ? write its two functions.

Ans. Formation of Salt Bridge:
A U-shaped glass tube filled with an inert electrolyte solution (like KCl or KNO_3) in agar-agar or gelatin is used to connect the two half-cells.

Functions:

1. Maintains electrical neutrality by allowing ions to move between the two half-cells.
2. Completes the electrical circuit for the flow of current.

Q10. a) State Faraday's first law of electrolysis.

b) What is the mass of silver deposited when 300 coulomb electricity is passed through a solution of $AgNO_3$? (Atomic mass of Ag = 108 u)

Ans. **a) Faraday's First Law of Electrolysis:**
The **mass of a substance deposited or liberated at an electrode is directly proportional to the quantity of electricity passed through the electrolyte.**

$$m = \frac{Q \times M}{nF}$$

b) Mass of Ag deposited

Given:

$Q = 300\,C,\ M = 108,\ n = 1,\ F = 96500\,C$

$$m = \frac{300 \times 108}{1 \times 96500}$$

$$m = \frac{32400}{96500} \approx 0.336\,g$$

Mass of Ag deposited ≈ 0.336 g.

17

Hydrogen and s-Block Elements

1. Hydrogen

Unique Position: Hydrogen shows similarities with both alkali metals (**Group 1**) and halogens (**Group 17**), making its position unique in the periodic table.

Isotopes: It has three isotopes: **Protium** $({}_1^1H)$ **Deuterium** $({}_1^2H$ or $D)$ and **Tritium** $({}_1^3H$ or $T)$

Tritium is radioactive.

Water and Ice: Water has a bent structure. Ice has an open cage-like structure due to hydrogen bonding, which is why ice is less dense than water.

Hydrogen Peroxide (H_2O_2): Acts as both an oxidizing and a reducing agent. Used as a bleaching agent and germicide.

Hydrogen and s-Block

A. Preparation of H_2O_2 (Laboratory Method):

$$BaO_2 \cdot 8H_2O(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2O_2(aq) + 8H_2O(l).$$

B. Solvay Process (Sodium Carbonate Manufacture):

$$NH_3 + H_2O + CO_2 \rightarrow NH_4HCO_3$$

$$NH_4HCO_3 + NaCl \rightarrow NaHCO_3 \downarrow + NH_4Cl$$

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2.$$

2. s-Block Elements: Alkali Metals (Group 1)

Elements: Li, Na, K, Rb, Cs, Fr

Properties: Highly reactive, soft metals, low ionization enthalpy. They impart characteristic colors to the flame (e.g., Na is yellow, K is violet).

Reactivity: They react vigorously with water to form hydroxides and hydrogen gas.

3. s-Block Elements: Alkaline Earth Metals (Group 2)

Elements: Be, Mg, Ca, Sr, Ba, Ra

Properties: Harder than alkali metals, higher melting points, and usually form **+2** oxidation states.

Biological Role: **Mg^{2+}** is essential for chlorophyll in plants; **Ca^{2+}** is vital for bones, teeth, and blood clotting.

4. Important Compounds

Sodium Hydroxide (NaOH): Caustic soda, used in soap manufacture.

Plaster of Paris: $CaSO_4 \cdot \frac{1}{2}H_2O$ used in building materials and setting fractured bones.

Top 10 Expected Questions

1. Explain why hydrogen is placed separately in the periodic table.

Ans. Hydrogen is often called the "rogue" element because it doesn't fit perfectly into any single group. While it sits at the top of Group 1, it is kept separate for these 3 main reasons:

1. Dual Nature (Like Metals and Non-metals)

Hydrogen behaves like two different groups at once:

- Like Alkali Metals (Group 1): It has 1 electron in its outer shell and can lose it to form a positive ion (H^+).
- Like Halogens (Group 17): It is a non-metal and only needs 1 more electron to fill its shell. It can gain an electron to form a hydride ion (H^-).

2. Physical Properties

Unlike the other members of Group 1 (which are soft, reactive solids), hydrogen is a gas at room temperature. It does not conduct electricity or heat, which is the opposite of how metals behave.

3. Unique Size

When hydrogen loses its electron, it becomes a bare proton (H^+). This ion is extremely small—much smaller than any other atom's ion. Because of this high "charge density," it behaves differently in chemical reactions compared to larger atoms.

2. Compare the three isotopes of hydrogen based on their mass and radioactivity.

Ans.

Feature	Protium (11H)	Deuterium (12H)	Tritium (13H)
Common Name	Ordinary Hydrogen	Heavy Hydrogen	Radioactive Hydrogen
Neutrons	0	1	2
Mass Number	1 (Lightest)	2	3 (Heaviest)

Radioactivity	Non-radioactive	Non-radioactive	Radioactive

3. Why does ice float on water? Explain with reference to its structure.

Ans. 1. Open Cage-Like Structure

In liquid water, molecules are constantly moving and are close together. However, as water freezes into ice, each water molecule forms four hydrogen bonds with neighboring molecules. These bonds lock the molecules into a rigid, hexagonal cage-like structure (lattice).

2. Increased Volume

Because of this open cage-like arrangement, the molecules in ice are actually pushed further apart than they were in the liquid state. This causes the same number of molecules to occupy a larger space, meaning the volume increases.

3. Lower Density

Since the volume increases while the mass remains the same, the density of ice becomes lower than that of liquid water.

$$Density = \frac{Mass}{Volume}$$

Because the density of ice is about 9% less than that of liquid water, it remains on the surface

4. Write the chemical equation for the preparation of H_2O_2 from barium peroxide.

Ans. The Chemical Equation

The reaction is as follows:

$$BaO_2 \cdot 8H_2O(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2O_2(aq) + 8H_2O(l)$$

Key Points

Hydrated Barium Peroxide: We use the hydrated form ($BaO_2 \cdot 8H_2O$) because anhydrous barium peroxide forms a protective layer of barium sulfate ($BaSO_4$) that stops the reaction.

Precipitation: The $BaSO_4$ formed is an insoluble white precipitate, which can be easily removed by filtration, leaving behind the H_2O_2 solution.

Temperature: The reaction is kept ice-cold to prevent the H_2O_2 from decomposing

5. What are the characteristic flame colors of Sodium and Potassium?

Ans. 1. Sodium (Na)

Flame Color: Golden Yellow

Reason: Sodium emits a very strong and persistent yellow light (specifically the "Sodium D-lines") which can often mask other colors if the sample is contaminated.

2. Potassium (K)

Flame Color: Lilac (Pale Violet)

Observation Tip: Because the lilac color is faint, it is often viewed through a blue cobalt glass. The glass filters out any yellow light from sodium impurities, making the violet flame easier to see.

6. Explain the diagonal relationship between Lithium and Magnesium.

Ans. 1. Reason for Similarity

The main reason is the similarity in their ionic sizes and polarizing power (charge/radius ratio).

Ionic Radii: Li^+ (76 pm) and Mg^{2+} (72 pm) are very close in size.

Electronegativity: Both have similar values $(Li \approx 1.0, Mg \approx 1.2)$

2. Key Similarities

- Hardness: Both Li and Mg are harder and have higher melting points than other elements in their respective groups.
- Reaction with Nitrogen: Both react directly with nitrogen to form nitrides (Li_3N and Mg_3N_2) Other alkali metals do not do this.
- Oxides: Both form only "normal" oxides (Li_2O and MgO) when heated in air, instead of peroxides or superoxides.
- Carbonates: Their carbonates (Li_2CO_3 and $MgCO_3$) are unstable and decompose easily on heating to release CO_2
- Solubility: Their fluorides, phosphates, and carbonates are relatively insoluble in water compared to other Group 1 elements.

3. Comparison Summary

Property	Lithium (Li)	Magnesium (Mg)
Ionic Radius	76 pm	72 pm
Nitride Formation	Forms Li_3N	Forms Mg_3N_2
Carbonate Stability	Decomposes on heating	Decomposes on heating

7. Write the chemical equations for the Solvay process to manufacture Na_2CO_3

Ans. The Solvay Process (Summary)

Goal: To manufacture Sodium Carbonate (Na_2CO_3).

The 3 Main Steps:

1. **Carbonation:**

$$NH_3 + H_2O + CO_2 + NaCl \rightarrow NaHCO_3 \downarrow + NH_4Cl$$

(Ammonia, water, and CO_2 react with brine to precipitate Sodium Bicarbonate).

2. **Calcination (Heating):**

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2 \uparrow$$

(Sodium Bicarbonate is heated to produce Sodium Carbonate).

3. **Ammonia Recovery:**

$$2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$$

(Ammonia is recovered to be reused in Step 1).

8. Why are alkali metals kept under kerosene?

Ans. 1. High Reactivity

Alkali metals are extremely reactive. If left in the open, they react vigorously with the oxygen (O_2) and moisture (H_2O) present in the air.

2. Prevention of Oxidation

When exposed to air, they quickly form a layer of oxides and hydroxides on their surface, causing them to lose thei

r metallic luster (tarnish).

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

3. Fire Hazard

The reaction with water/moisture is exothermic (releases heat) and produces hydrogen gas, which can catch fire instantly.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2 \uparrow + \text{Heat}$$

9. Explain the biological importance of Mg^{2+} and Ca^{2+} ions.

Ans. Biological Importance (Short)

Magnesium (Mg^{2+}) - Power & Plants

- Photosynthesis: Central atom in Chlorophyll (essential for plants to make food).
- Energy: Necessary for all enzymes that use or travel with ATP (energy currency).
- DNA/RNA: Helps in the synthesis and stability of genetic material.

Calcium (Ca^{2+}) - Structure & Signals

- Bones & Teeth: Provides structural strength (99% of body calcium is in bones).
- Blood Clotting: Essential for stopping bleeding after an injury.
- Muscle & Nerve: Triggers muscle contraction and helps in sending nerve signals.

10. Describe the use of Heavy Water (D_2O) in nuclear reactors.

Ans. **1. As a Moderator**

The main role of heavy water is to slow down fast-moving neutrons produced during nuclear fission.

- Fission produces "fast neutrons," but for a sustained chain reaction with Uranium-235, "slow neutrons" (thermal neutrons) are required.
- D_2O is an excellent moderator because it slows neutrons effectively without absorbing them, unlike ordinary water (H_2O).

2. As a Coolant

Heavy water acts as a heat transfer medium.

- It absorbs the massive amount of thermal energy (heat) generated in the reactor core.
- This heat is then transferred to a secondary system to produce steam, which drives turbines to generate electricity.

18

General Characteristics of p-Block Elements

1. Introduction to p-Block

Position: Elements of Groups **13, 14, 15, 16, 17**, and **18**.

Electronic Configuration: The general valence shell electronic configuration is $ns^2\ np^{1-6}$ (except for Helium which is $1s^2$).

Nature: This block is unique as it contains metals, non-metals, and metalloids.

2. Atomic and Physical Trends

Atomic Radius: Decreases across a period (**left to right**) and increases down a group.

Ionization Enthalpy: Increases across a period and decreases down a group.

Electron Gain Enthalpy: Generally becomes more negative across a period.

Note: Fluorine has a lower electron gain enthalpy than Chlorine due to its extremely small size and high inter-electronic repulsion.

3. Chemical Properties and Anomalies

Inert Pair Effect: In heavier elements (like **Tl, Pb, Bi**), the s-electrons of the valence shell do not participate in bonding. This makes lower oxidation states more stable (e.g., Pb^{2+} is more stable than Pb^{4+}).

Catenation: The ability of an element to form long chains or rings by linking with its own atoms. Carbon shows the maximum property of catenation.

Anomalous Behavior of Second Period Elements: The first element of each group (like **N, O, F**) shows different properties due to:

1. Small size and high electronegativity.
2. Absence of vacant **d-orbitals**.

Top 10 Expected Questions

1. Define the p-block and give the general electronic configuration of these elements.

Ans. 1. Definition

Elements in which the last electron (valence electron) enters any of the three p-orbitals (px, py, or pz) of their outermost shell are called p-block elements. This block includes metals, metalloids, and non-metals.

2. General Electronic Configuration

The general valence shell electronic configuration for p-block elements is:

$ns^2\ np^{1-6}$

(Where n is the principal quantum number or the period number, ranging from 2 to 7).

2. Why is Chlorine's electron gain enthalpy higher than that of Fluorine?

Ans. 1. Atomic Size

Fluorine (F) is a very small atom compared to Chlorine (Cl).

In Fluorine, the 2p subshell is very compact.

In Chlorine, the 3p subshell is larger and has more space.

2. Inter-electronic Repulsion

When an extra electron is added to the small 2p orbital of Fluorine, it faces strong repulsion from the electrons already present there. This "crowding" makes it harder for Fluorine to welcome a new electron.

3. Chlorine's Advantage

In Chlorine, the 3p orbital is spread over a larger region of space. The incoming electron experiences much less repulsion from the existing electrons, allowing the nucleus to hold it more effectively.

3. Explain the Inert Pair Effect with an example of Group 14 elements.

Ans. 1. The Cause

In heavier elements (like those in Periods 5 and 6), the d and f orbitals provide poor shielding. This allows the nucleus to exert a very strong attraction on the outermost s-electrons, making them reluctant to be lost or shared.

2. Example: Group 14 Elements

Group 14 elements (Carbon family) have the general configuration ns^2 np^2. They can show two oxidation states: +4 (using all 4 electrons) and +2 (using only the 2 p-electrons).

- Lighter Elements (C, Si): The +4 state is very stable.
- Heavier Elements (Sn, Pb): The +2 state becomes more stable than the +4 state due to the inert pair effect.

3. Specific Case of Lead (Pb)

For Lead (Pb), the $6s^2$ electrons are so "inert" that:

- Pb^{2+} is very stable.
- Pb^{4+} is unstable and acts as a strong oxidizing agent (it wants to gain 2 electrons to return to the stable +2 state).

4. Why do the first elements of each group in the p-block show anomalous behavior?

Ans. Anomalous Behavior (Short Summary)

1. Tiny Size: They are the smallest in their group, leading to high charge density.
2. No d-orbitals: Because they are in the 2nd Period, they lack d-orbitals and cannot expand their octet (e.g., Nitrogen forms NF_3, but Phosphorus can form PF_5).
3. High Grip: They have the highest Electronegativity and Ionization Enthalpy in their group.
4. π Bonding: Their small size allows them to form strong multiple bonds (p\pi - p\pi), like C=C or N ≡ N, which heavier elements cannot do easily.

5. Compare the stability of +2 and +4 oxidation states in Tin (Sn) and Lead (Pb).

Ans.

Element	Most Stable State	Relative Stability
Tin (Sn)	+4	$Sn^{4+} > Sn^{2+}$
Lead (Pb)	+2	$Pb^{2+} > Pb^{4+}$

6. Explain why Oxygen exists as a gas (O_2) while Sulphur exists as a solid (S8).

Ans. Oxygen (O_2) vs. Sulphur (S_8)

1. Bonding: Oxygen is small and forms strong double bonds (O=O). Sulphur is larger and prefers single bonds (S-S).
2. Structure: Oxygen exists as tiny diatomic molecules (O_2). Sulphur forms large, 8-atom puckered rings (S_8) in a crown shape.
3. Molecular Mass: O_2 has a low mass (32), while S_8 has a high mass (256).
4. Forces: S_8 has much stronger Van der Waals forces due to its larger size and mass, making it a solid. O_2 has weak forces, making it a gas.

7. What is Catenation? Why is it most prominent in Carbon?

Ans. Definition: Catenation is the unique ability of an element to form long chains or rings by bonding with its own atoms through covalent bonds.

Why is it most prominent in Carbon? Carbon shows the highest degree of catenation in the periodic table due to two main reasons:

1. High Bond Enthalpy: The Carbon-Carbon (C-C) bond is extremely strong and stable (348 kJ/mol). This allows long chains to stay together without breaking easily.
2. Small Atomic Size: Because Carbon is small, the shared pair of electrons is held strongly by the nucleus, making the covalent bonds very robust.

3. Tetravalency: With four valence electrons, Carbon can bond in multiple directions (linear, branched, or cyclic), allowing for a vast variety of structures.

8. Which is more metallic: Nitrogen or Bismuth? Justify based on group trends.

Ans. Justification Based on Group Trends

In Group 15 (the Nitrogen family), the metallic character increases as we move down the group.

1. Atomic Size: As we go down from Nitrogen to Bismuth, new electronic shells are added, making the atomic size significantly larger.
2. Ionization Enthalpy: Because the outer electrons are farther from the nucleus in Bismuth, they are held less tightly. This means Bismuth has a lower ionization enthalpy and can lose electrons more easily than Nitrogen.
3. Electronegativity: Nitrogen is highly electronegative (a typical non-metal), while Bismuth has a much lower electronegativity, allowing it to show metallic properties.

9. Explain the trend of electronegativity in the p-block across a period and down a group.

Ans. Electronegativity Trends (Short Summary)

1. **Across a Period (Left to Right): * Trend: Increases.**

- Reason: Atomic size decreases and nuclear charge increases, pulling shared electrons more strongly.
- Highest: Fluorine (F) is the most electronegative.

2. **Down a Group (Top to Bottom): * Trend: Decreases.**

- Reason: Atomic size increases as new shells are added; the nucleus is farther from the shared electrons.
- Lowest: Elements at the bottom (like Bismuth or Thallium) have the lowest values.

10. Give one industrial use of Nitrogen and Phosphorus compounds.

Ans. Here are the primary industrial uses for Nitrogen and Ammonia-based compounds and Phosphorus compounds:

1. Nitrogen Compounds (Ammonia - NH_3)

- Industrial Use: Fertilizer Production.
- Detail: The most significant use of nitrogen is in the Haber Process to produce Ammonia. About 80% of this ammonia is used to manufacture nitrogenous fertilizers like Urea, Ammonium Nitrate, and Ammonium Phosphate, which are essential for global food production.

2. Phosphorus Compounds (Phosphoric Acid - H_3PO_4)

- Industrial Use: Detergents and Water Treatment.
- Detail: Phosphorus compounds, specifically Sodium Tripolyphosphate (STPP), are used in detergents to soften hard water. Additionally, Phosphoric acid is widely used in the food industry as an acidulant in soft drinks (like colas) to provide a tangy taste and prevent the growth of bacteria.

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p-Block Elements & Their Compounds - II

1. Group 16: The Oxygen Family (Chalcogens)

Elements: **O, S, Se, Te, Po**

Oxides Classification:

Acidic: React with water to give acids (e.g., **SO_2, CO_2**).

Basic: React with water to give bases (e.g., **Na_2O, CaO**).

Amphoteric: Show both acidic and basic properties (e.g., **Al_2O_3, ZnO**)

Ozone (O_3): An allotrope of oxygen. It is a powerful oxidizing agent and has a bent structure with a bond angle of **117°**

$$2Hg + O_3 \rightarrow Hg_2O + O_2$$

Ozone Reactions:

Tailing of Mercury: $2Hg + O_3 \rightarrow Hg_2O + O_2$ (This causes Mercury to stick to glass).

Oxidizing Property: $PbS(s) + 4O_3(g) \rightarrow PbSO_4(s) + 4O_2(g)$.

Sulphuric Acid (H_2SO_4): Manufactured by the **Contact Process**. It acts as a strong oxidizing agent and a powerful dehydrating agent.

$S + O_2 \rightarrow SO_2$

$2SO_2 + O_2 \xrightarrow{V_2O_5} 2SO_3$ (Exothermic and reversible)

$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$ (Oleum)

$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$

2. Group 17: The Halogens

Elements: **F, Cl, Br, I, At**

Hydrogen Halides: Boiling points follow the order **HF > HI > HBr > HCl. HF** has the highest boiling point due to strong hydrogen bonding.

Interhalogen Compounds: Formed when two different halogens react (e.g., **ClF, BrF_3, IF_7**). They are more reactive than pure halogens.

Oxoacids of Chlorine: Strength increases with the oxidation state of Chlorine: **HClO < $HClO_2$ < $HClO_3$ < $HClO_4$**

3. Group 18: Noble Gases

Elements: **He, Ne, Ar, Kr, Xe, Rn**

Characteristics: Completely filled valence shells **(ns^2 np^6),** making them chemically inert.

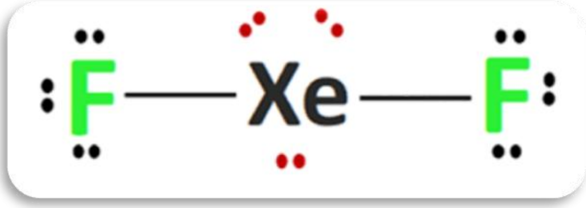
Xenon Compounds: Xenon reacts with fluorine and oxygen to form compounds like **XeF_2**(Linear), **XeF_4** (Square Planar), and **XeO_3** (Pyramidal).

Xenon Compounds (Preparation):

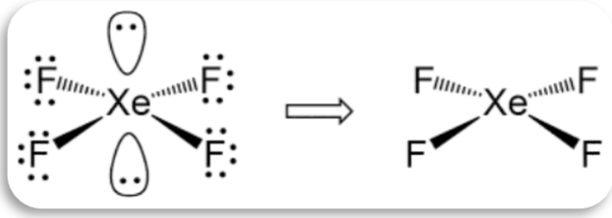
$$Xe(g) + F_2(g) \xrightarrow{673K,1bar} XeF_2(s)$$

$$Xe(g) + 2F_2(g) \xrightarrow{873K,7bar} XeF_4(s)$$

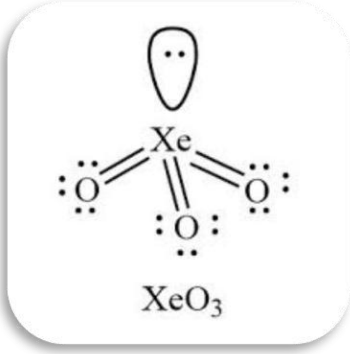
XeF_2(Linear)



XeF_4 (Square Planar)



XeO_3 (Pyramidal)



Top 10 Expected Questions

1. Classify the following oxides as acidic, basic, or amphoteric: SO_3, Al_2O_3, CaO, CO_2, ZnO

Ans.

Oxide	Type	Reason
CaO	Basic	Metal oxide that reacts with water to form a base, $Ca(OH)_2$
SO_3	Acidic	Non-metal oxide that reacts with water to form Sulphuric acid, H_2SO_4
CO_3	Acidic	Non-metal oxide that reacts with water to form Carbonic acid, H_2CO_3
Al_2O_3	Amphoteric	Reacts with both acids and bases to form salts and water.
ZnO	Amphoteric	Reacts with both acids and bases to form salts and water.

2. Explain the "Tailing of Mercury" when exposed to ozone.

Ans. Reaction: Ozone (O_3) oxidizes Mercury (Hg) to Mercurous Oxide (Hg_2O).

$2HG + O_3 \rightarrow Hg_2O + O_2$

Effect: The Hg_2O dissolves in the mercury, making it stick to the glass instead of moving as a clean drop.

Appearance: The mercury loses its curved shape (meniscus) and leaves a "tail" or trail on the glass surface.

Fix: Shaking it with water removes the oxide and restores the mercury to its normal state.

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

Ans. 1. Hydrogen Bonding in H_2O

Oxygen (O) is highly electronegative and small in size. This allows water molecules to form strong intermolecular hydrogen bonds.

- These strong forces hold the H_2O molecules closely together, making it a liquid at room temperature.

2. Van der Waals Forces in H_2S

Sulphur (S) is larger and has low electronegativity compared to Oxygen. It cannot form hydrogen bonds.

- Instead, H_2S molecules are held together by much weaker Van der Waals forces (dipole-dipole interactions).
- Because these forces are weak, the molecules stay far apart, making H_2S a gas.

4. Describe the steps involved in the manufacture of Sulphuric acid by the Contact Process.

Ans. Goal: To manufacture Sulphuric Acid (H_2SO_4).

1. Burning: Sulphur is burnt in air to get Sulphur Dioxide.

$$S + O_2 \rightarrow SO_2$$

2. Oxidation: SO_2 reacts with O_2 using a catalyst (V_2O_5) to get Sulphur Trioxide.

$$2SO_2 + O_2 \xrightarrow{V_2O_5} 2SO_3$$

3. Absorption: SO_3 is dissolved in conc. H_2SO_4 to form Oleum ($H_2S_2O_7$).

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

4. Dilution: Oleum is mixed with water to get Sulphuric Acid.

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$

5. Why is HF a liquid whereas other hydrogen halides are gases?

Ans. 1. Hydrogen Bonding in HF

Fluorine (F) is the most electronegative element and has a very small atomic size. This allows HF molecules to form strong intermolecular hydrogen bonds.

- These strong attractive forces hold the HF molecules closely together in an associated form, making it a liquid at room temperature (boiling point: 293 K).

2. Van der Waals Forces in Others

Other halogens (Cl, Br, I) have lower electronegativity and larger atomic sizes. They are unable to form hydrogen bonds.

- Their molecules are held together only by weaker Van der Waals forces (specifically dipole-dipole interactions).
- Because these forces are weak, the molecules stay far apart, making HCl, HBr, and HI gases at room temperature.

6. Arrange the oxoacids of chlorine in increasing order of their acidic strength and justify your answer.

Ans. The acidic strength increases as the oxidation state of the central Chlorine (Cl) atom increases:

1. **Oxidation States:**
   - HOCl (Hypochlorous acid): +1
   - $HClO_2$ (Chlorous acid): +3
   - $HClO_3$ (Chloric acid): +5
   - $HClO_4$ (Perchloric acid): +7
2. **Electronegativity Effect:** As the number of oxygen atoms increases, the oxidation state of Chlorine becomes more positive. This makes the Cl atom more electron-withdrawing, which weakens the O-H bond, making it easier to release a proton (H^+).
3. **Resonance Stability:** After losing a proton, the resulting anion ($ClO4^-$, $ClO3^-$, etc.) becomes more stable. The more oxygen atoms there are, the more the negative charge can spread out (delocalize) through resonance.
   - $ClO4^-$ is the most stable anion because the charge is spread over four oxygen atoms.

○ OCl^- is the least stable.

7. What are Interhalogen compounds? Why are they more reactive than halogens?

Ans. Definition: Interhalogen compounds are molecules formed when two different halogen atoms (X and X') react with each other. Their general formula is XX'n, where n = 1, 3, 5, or 7.

- Example: ICl (Iodine monochloride), ClF_3 (Chlorine trifluoride), IF_7 (Iodine heptafluoride).

Interhalogen compounds are generally more reactive than the pure halogens (except for Fluorine).

1. Bond Polarity: In a pure halogen (like Cl-Cl), the bond is non-polar. In an interhalogen (like I-Cl), the bond is polar because the two atoms have different electronegativities.
2. Bond Strength: The X-X' bond in interhalogens is weaker than the X-X bond in pure halogens.
   - *Reason:* The overlap between orbitals of two different-sized atoms is less effective than between two atoms of the same size.
3. Easier Cleavage: Because the bond is weaker and polar, it breaks much more easily during a chemical reaction, making the compound highly reactive.

8. Explain the structure of XeF_4 and XeO_3 based on VSEPR theory.

Ans. 1. Structure of XeF_4 (Xenon Tetrafluoride)

- Valence Electrons: Xe has 8 electrons. It forms 4 bonds with Fluorine (F).
- Electron Pairs:
  - Bond Pairs (bp): 4
  - Lone Pairs (lp): 2 (since 8 - 4 = 4 electrons left)
  - Total Pairs: 4 + 2 = 6 (Octahedral geometry)
- Shape: Square Planar.

- Reason: To minimize repulsion, the two lone pairs occupy opposite axial positions, leaving the four Fluorine atoms in a square plane.

2. Structure of XeO_3 (Xenon Trioxide)

- Valence Electrons: Xe has 8 electrons. It forms 3 double bonds with Oxygen (O).
- Electron Pairs:
  - Bond Pairs (bp): 3 (Each double bond is counted as one bonding center)
  - Lone Pairs (lp): 1 (since 8 - 6 = 2 electrons left)
  - Total Pairs: 3 + 1 = 4 (Tetrahedral geometry)
- Shape: Pyramidal.
- Reason: The three bonding pairs and one lone pair arrange themselves tetrahedrally, but the presence of the lone pair pushes the bonds down, resulting in a pyramid shape.

9. Mention two important uses of Noble gases.

Ans. 1. Helium (He)

- Cryogenic Agent: Helium has the lowest boiling point of any element. It is used as a liquid coolant to provide the extremely low temperatures required for superconducting magnets in MRI (Magnetic Resonance Imaging) machines.
- Balloons and Airships: Because it is very light and non-flammable (unlike hydrogen), it is used for filling weather balloons and decorative balloons.

2. Argon (Ar)

- Inert Atmosphere: Argon is used to provide an unreactive environment in high-temperature industrial processes, such as arc welding and the production of high-quality stainless steel, to prevent the metals from oxidizing.
- Electric Bulbs: It is filled in incandescent light bulbs to prevent the tungsten filament from burning away at high temperatures.

10. What is the environmental impact of Chloro Fluoro Carbons (CFCs)?

Ans. 1. Ozone Layer Depletion

This is the most significant impact. When CFCs reach the stratosphere, ultraviolet (UV) radiation breaks them down, releasing Chlorine atoms.

- Chain Reaction: A single chlorine atom can destroy thousands of ozone (O_3) molecules.
- Result: This creates "ozone holes," allowing harmful UV-B radiation to reach the Earth's surface, increasing risks of skin cancer, cataracts, and damage to marine life.

2. Global Warming (Greenhouse Effect)

CFCs are incredibly potent Greenhouse Gases.

- Heat Trapping: They are much more effective at trapping heat in the atmosphere than Carbon Dioxide (CO_2).
- Result: Even in small amounts, they contribute significantly to rising global temperatures and climate change.

21

d-Block and f-Block Elements

1. The d-Block (Transition Elements)

Definition: Elements in which the last electron enters the (n-1)d subshell. They have the general electronic configuration **$(n-1)d^{1-10}$ ns^{1-2}**

General Characteristics:

Variable Oxidation States: They show many oxidation states because of the small energy gap between **(n-1)d** and **ns** orbitals (e.g., Manganese shows from **+2 to +7).**

Colored Ions: Most transition metal ions are colored due to d-d transitions.

Catalytic Properties: They act as catalysts because of their ability to show multiple oxidation states (e.g., **Fe** in Haber's process).

Magnetic Properties: Many ions are paramagnetic due to the presence of unpaired electrons.

2. Important Compounds

A. Potassium Dichromate ($K_2Cr_2O_7$):

Preparation: From Chromite ore **($FeCr_2O_4$)**

Structure: The dichromate ion **($Cr_2O_7^{2-}$)** consists of two tetrahedra sharing one oxygen atom.

Reaction with Base: It changes to Chromate in alkaline medium.

$$Cr_2O_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$$

B. Potassium Permanganate ($KMnO_4$):

Preparation: From Pyrolusite ore (MnO_2).

Structure: The Permanganate ion (MnO_4^-) is tetrahedral.

Key Reaction (Oxidizing Agent in Acidic Medium): $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

3. The f-Block (Inner Transition Elements)

Lanthanoids: **14** elements from Cerium **(58)** to Lutetium **(71).**

Lanthanoid Contraction: The steady decrease in atomic and ionic radii with increasing atomic number. This is due to poor shielding of **4f** electrons.

Actinoids: **14** elements from Thorium **(90)** to Lawrencium **(103).** Most are radioactive

Top 10 Expected Questions

1. What is Lanthanoid contraction? Mention its consequences.

Ans. Definition:

As the atomic number increases from Lanthanum (La) to Lutetium (Lu), there is a steady and significant decrease in the size of atoms and M^{3+} ions. This gradual shrinkage is called Lanthanoid Contraction.

The Cause:

It occurs because of the poor shielding effect of 4f electrons. As the nuclear charge increases, the 4f electrons fail to protect the outer electrons from the nucleus, pulling them closer and shrinking the atom.

Consequences

Similarity in 4d and 5d Series: Due to the contraction, the atomic radii of the second (4d) and third (5d) transition series become almost identical (e.g., Zirconium (Zr) and Hafnium (Hf) have nearly the same size).

Difficulty in Separation: Because their sizes and chemical properties are so similar, it is extremely difficult to separate Lanthanoids from one another in their pure state.

Basic Character of Hydroxides: The basic strength of hydroxides decreases from $La(OH)_3$ to $Lu(OH)_3$ As the size of the M^{3+} ion decreases, the covalent character of the M-OH bond increases, making it harder to release OH^- ions.

2. Why do transition elements show variable oxidation states?

Ans. Reason: There is a very small energy difference between the (n-1)d and ns orbitals (e.g., 3d and 4s).

Participation: Electrons from both these orbitals can take part in bond formation.

Trend:

- Start of series: Fewer oxidation states (Sc is +3).
- Middle of series: Maximum variety (Mn shows +2 to +7).
- End of series: Fewer states as d-orbitals become completely filled.

3. Calculate the magnetic moment of Fe^{2+} (At. No. 26).

Ans. Step-by-Step Calculation

1. Electronic Configuration of Fe (At. No. 26):

[Ar] $3d^6$ $4s^2$

2. Electronic Configuration of Fe^{2+}:

The two electrons are lost from the 4s orbital.

[Ar] $3d^6$

3. Number of Unpaired Electrons (n):

In the $3d^6$ orbital, electrons fill as: (↑↓) (↑) (↑) (↑) (↑)

There are n = 4 unpaired electrons.

4. Magnetic Moment Formula:

$$\mu = \sqrt{n(n+2)}\ \text{BM (Bohr Magnetons)}$$

Calculation:

$$\mu = \sqrt{4(4+2)}$$

$$\mu = \sqrt{4 \times 6} = \sqrt{24}$$

$$\mu \approx \mathbf{4.89}\ \text{BM}$$

4. Explain why Cu^+ is colorless but Cu^{2+} is colored.

Ans. **1. Copper (I) Ion: Cu^+ (Colorless)**

- Electronic Configuration: Cu is [Ar] $3d^{10}\ 4s^1$. For Cu^+, it loses one electron from 4s, becoming [Ar] $3d^{10}$
- Reason: The 3d orbital is completely filled (10 electrons). Since there are no empty spaces or unpaired electrons, d-d transitions are impossible. Therefore, Cu^+ compounds are colorless.

2. Copper (II) Ion: Cu^{2+} (Colored/Blue)

- Electronic Configuration: For Cu^{2+}, it loses two electrons (one from 4s and one from 3d), becoming [Ar] $3d^9$
- Reason: The 3d orbital is incompletely filled. It has one unpaired electron. This electron can absorb visible light and jump from a lower-energy d-orbital to a higher-energy d-orbital (d-d transition). The reflected or transmitted light gives Cu^{2+} its characteristic blue color.

5. Describe the preparation of $KMnO_4$ from Pyrolusite ore.

Ans. From Pyrolusite Ore (MnO_2):

1. **Step 1: Fusion**

MnO_2 (Black) is heated with KOH and air (O_2) to form Potassium Manganate (Green).

$$2MnO_2 + 4KOH + O_2 \rightarrow \mathbf{2K_2MnO_4}(\text{Green}) + 2H_2O$$

2. **Step 2: Oxidation**

The green K_2MnO_4 is oxidized to Potassium Permanganate (Purple) using Chlorine or Electrolysis.

$$2K_2MnO_4 + Cl_2 \rightarrow \mathbf{2KMnO_4}(\text{Purple}) + 2KCl$$

6. Write the ionic equation for the reaction between acidic $K_2Cr_2O_7$ and I^-

Ans. 1. **Reduction Half-Reaction (Dichromate):**

The dichromate ion ($Cr2O^{2-}{}_7$) gains electrons to become Chromium (III):

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

2. **Oxidation Half-Reaction (Iodide):**

Iodide ions (I^-) lose electrons to form Iodine (I^2):

$$6I^- \rightarrow 3I_2 + 6e^-$$

3. **Overall Balanced Ionic Equation**

By adding the two half-reactions together, we get the final balanced equation:

$$\mathbf{Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O}$$

7. Compare the properties of Lanthanoids and Actinoids.

Ans. Lanthanoids vs. Actinoids

The primary difference lies in the filling of 4f orbitals (Lanthanoids) versus 5f orbitals (Actinoids).

Feature	Lanthanoids (4f series)	Actinoids (5f series)
Radioactivity	Except Promethium, they are non-radioactive.	All elements are radioactive.
Oxidation States	Mainly +3 (rarely +2 or +4).	Show a wider range (+3, +4, +5, +6, +7).
Complex Formation	Less tendency to form complexes.	High tendency to form complexes.
Magnetic Properties	Magnetic properties are easier to explain.	Magnetic properties are complex and difficult to explain.
Basicity	Less basic.	More basic than lanthanoids.

Oxoions	Do not form oxoions.	Form oxoions like UO^{2+}_2, PuO^{2+}_2

8. Transition metals and their compounds are good catalysts. Why?

Ans. Variable Oxidation States: They can easily change their oxidation state to form unstable intermediates, which lowers the activation energy of a reaction.

Adsorption (Surface Area): They provide a solid surface where reactant molecules can stick (adsorb). This increases the concentration of reactants and weakens their bonds, making the reaction faster.

Empty d-orbitals: They have vacant d-orbitals that can accept electrons from reactant molecules to form temporary bonds.

9. What happens when OH^- ions are added to a solution of $K_2Cr_2O_7$?

Ans. Observation: The orange solution of Potassium Dichromate turns yellow.

Chemical Change: The Dichromate ion ($Cr_2O^{2-}_7$) converts into the Chromate ion (CrO^{2-}_4).

$$Cr_2O_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$$

10. Describe the general electronic configuration of f-block elements.

Ans. The general electronic configuration of f-block elements (Inner Transition Elements) is:

$(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$

Key Breakdown

- $(n-2)f^{1-14}$: Electrons fill the third-to-last shell (f-orbital).
- $(n-1)d^{0-1}$: The second-to-last shell has 0 or 1 electron.
- ns^2: The outermost shell always has 2 electrons.

22

Coordination Compounds

1. Introduction

Coordination compounds are those in which a central metal atom or ion is surrounded by a group of ions or molecules. They retain their identity even in a solution.

Example: Chlorophyll (Magnesium complex) and Hemoglobin (Iron complex).

2. Key Terms (Important Definitions)

Coordination Entity: A central metal atom or ion bonded to a fixed number of ions or molecules. **Example:** In **$[CoCl_3(NH_3)_3]$,** the entire bracketed part is the entity.

Central Atom/Ion: The atom to which a fixed number of ions/groups are bound in a definite geometrical arrangement.

Ligands: The ions or molecules bound to the central atom. They act as **Lewis Bases** by donating electron pairs.

Unidentate: One donor atom **(e.g., Cl^-, NH_3)**

Didentate: Two donor atoms **(e.g., Ethylene diamine 'en').**

Polydentate: Many donor atoms **(e.g., EDTA).**

Coordination Number (CN): The total number of ligand donor atoms to which the metal is directly bonded.

Example: In **$[PtCl_6]^{2-}$, CN is 6.**

3. Werner's Theory

Alfred Werner proposed that metals show two types of valencies:

1. **Primary Valency:** Ionizable, corresponds to the oxidation state.
2. **Secondary Valency:** Non-ionizable, corresponds to the coordination number and has a fixed spatial arrangement (**geometry**).

4. Nomenclature (IUPAC Rules)

1. Name the cation first, then the anion.
2. Ligands are named in alphabetical order.
3. For anionic complexes, the metal name ends in **'-ate'** (e.g., Ferrate, Platinate).

Example: **$K_4[Fe(CN)_6]$** is named as **Potassium hexacyanoferrate (II)**.

5. Valence Bond Theory (VBT)

Explains the bonding, geometry, and magnetic behavior of complexes:

sp^3 Hybridization: Tetrahedral geometry **(e.g., $[NiCl_4]^{2-}$).**

dsp^2 Hybridization: Square Planar geometry **(e.g., $[Ni(CN)_4]^{2-}$).**

d^2sp^3 or sp^3d^2 Hybridization: Octahedral geometry **(e.g., $[Fe(CN)_6]^{3-}$).**

6. Crystal Field Theory (CFT)

Describes the splitting of d-orbitals when ligands approach the metal.

Octahedral Splitting: The **5 d**-orbitals split into two sets: **t_2g** (lower energy) and **eg** (higher energy).

Spectrochemical Series: Arranges ligands by their splitting power.

Strong Field (CN^-): Causes large splitting, forces electron pairing (**Low Spin**).

Weak Field (F^-): Causes small splitting, no pairing (**High Spin**).

7. Isomerism

Structural: Linkage (different donor atoms), Ionization, and Coordination isomerism.

Stereo-isomerism:

Geometrical: **Cis** (same side) and **Trans** (opposite sides).

Optical: Mirror images that are non-superimposable (Enantiomers).

Top 10 Expected Questions

1. Define Ligand and classify them based on the number of donor atoms.

Ans. Definition:

A ligand is an atom, ion, or molecule that donates a pair of electrons to a central metal atom to form a coordinate bond.

Classification (Based on Donor Atoms)

Type	No. of Donor Atoms	Example
Unidentate	1	NH3 (Ammonia), Cl-
Bidentate	2	Ethane-1,2-diamine (en), Oxalate ($C2O^{2-}{}_{4}$)
Polydentate	3 or more	EDTA (Hexadentate - 6 atoms)
Ambidentate	2 (but uses only 1 at a time)	$NO^{-}{}_{2}$ (Nitro), SCN^{-} (Thiocyanato)

2. Explain Werner's theory of coordination compounds with an example.

Ans. Werner's Theory (Very Short)

Alfred Werner was the first to explain the bonding in coordination compounds. His theory is based on two types of valencies for a metal:

1. **Primary Valency (Ionizable):**
   - Corresponds to the Oxidation State of the metal.
   - It is satisfied by negative ions.
   - It is non-directional.
2. **Secondary Valency (Non-Ionizable):**
   - Corresponds to the Coordination Number (number of ligands).
   - It is satisfied by negative ions or neutral molecules (ligands).
   - It is directional, which gives the complex a specific shape (e.g., Octahedral).

Example: $CoCl_3 \cdot 6NH_3$

In the modern formula $[Co(NH_3)_6]Cl_3$:

- Secondary Valency: 6 (The 6 NH_3 molecules are directly bonded to Cobalt).
- Primary Valency: 3 (The 3 Cl^- ions satisfy the charge and ionize in water).

3. Write the IUPAC name for $[Co(NH_3)_5Cl]Cl_2$ and $K_3[Fe(CN)_6]$

Ans. 1. $[Co(NH_3)_5Cl]Cl_2$

- Ligands: 5 Ammine (NH_3) and 1 Chlorido (Cl^-).
- Central Metal: Cobalt (Oxidation state = +3).
- Counter Ion: Chloride.
- IUPAC Name: Pentaamminechloridocobalt(III) chloride

2. $K_3[Fe(CN)_6]$

- Counter Ion: Potassium.
- Ligands: 6 Cyanido (CN^-).
- Central Metal: Iron (Oxidation state = +3). Since the complex part is an anion, the metal name ends in -ate (Ferrate).
- IUPAC Name: Potassium hexacyanidoferrate(III)

4. Distinguish between a Homoleptic and a Heteroleptic complex.

Ans. 1. Homoleptic Complex

- Definition: The metal is bound to only one type of ligand.
- Example: $[Co(NH_3)_6]^{3+}$ (All 6 ligands are NH_3).

2. Heteroleptic Complex

- Definition: The metal is bound to more than one type of ligand.
- Example: $[Co(NH_3)_4Cl_2]^+$ (Contains two types: 4 NH_3 and 2 Cl^-).

5. What is a Chelating Ligand? Give an example like EDTA.

Ans. Definition: A chelating ligand is a bidentate or polydentate ligand that uses two or more donor atoms to bind to a single metal ion, forming a ring-like structure.

- Result: The complex formed is much more stable than those formed with unidentate ligands. This is called the Chelate Effect.

Example: EDTA (Ethylene Diamine Tetraacetate)

EDTA is a famous hexadentate chelating ligand because it has 6 donor atoms (2 Nitrogen and 4 Oxygen).

- How it works: It wraps around the metal ion like a "claw," forming five stable 5-membered rings.
- Use: It is used to treat heavy metal poisoning (like lead) by "trapping" the metal and allowing it to be flushed from the body.

6. Using VBT, explain why $[Ni(CN)_4]^{2-}$ is diamagnetic while $[NiCl_4]^{2-}$ is paramagnetic.

Ans. 1. $[Ni(CN)_4]^{2-}$ (Diamagnetic)

- Ligand: CN^- is a Strong Field Ligand.
- Action: It forces the two unpaired 3d electrons to pair up.
- Hybridization: dsp2 (Square Planar).
- Result: No unpaired electrons, so it is Diamagnetic.

2- showing paired 3d electrons and dsp2 hybridization]

2. **$[NiCl_4]^{2-}$ (Paramagnetic)**

- Ligand: Cl^- is a Weak Field Ligand.
- Action: It cannot pair the 3d electrons.
- Hybridization: sp^3 (Tetrahedral).
- Result: 2 unpaired electrons remain, so it is Paramagnetic.

7. Why is $[T(H_2O)_6]^{3+}$ violet in color?

Ans. Configuration: Titanium (Ti) has an atomic number of 22. In $[Ti(H_2O)_6]^{3+}$, it is in the +3 oxidation state, giving it a $3d^1$ configuration (one unpaired electron).

Crystal Field Splitting: When the six H_2O ligands approach, the five d-orbitals split into two energy levels: the lower t_2g and the higher eg. The single electron stays in the lower t_2g level.

Absorption of Light: This electron absorbs energy from the visible spectrum (specifically yellow-green light) and jumps to the higher eg level. This is the d-d transition.

Complementary Color: Since yellow-green light is absorbed, the complementary color—which is violet—is transmitted or reflected, making the complex appear violet.

8. Calculate the oxidation state of the central metal in $[Cr(en)_3]Cl_3$

Ans. To find the oxidation state (x) of Chromium (Cr):

1. **Identify Charges:**
   - en (ethylenediamine): Neutral ligand (Charge = 0)
   - Cl (chloride): Counter ion (Charge = -1)
2. **Equation:**

$x + 3(0) + 3(-1) = 0$

$x - 3 = 0$

$x = +3$

9. Explain the hybridization and geometry of $[Fe(CN)_6]^{3-}$

Ans. To explain $[Fe(CN)_6]^{3-}$ using Valence Bond Theory (VBT):

1. Oxidation State: Fe^{3+} (Atomic No. 26). Configuration: $3d^5$
2. Ligand Action: CN^- is a Strong Field Ligand. It forces the 5 unpaired electrons in the 3d orbital to pair up.
3. Result: Two 3d orbitals become empty and available for bonding.
4. Hybridization: d^2sp^3 (Inner orbital complex).
5. Geometry: Octahedral.

10. What is the difference between cis and trans isomers? Draw the structure of $[Pt(NaH_3)_2Cl_2]$

Ans.

Feature	Cis-Isomer	Trans-Isomer
Arrangement	Identical ligands are adjacent (next to each other).	Identical ligands are opposite to each other.
Bond Angle	Typically 90° (in square planar).	Typically 180° (in square planar).
Dipole Moment	Usually has a non-zero dipole moment.	Usually has zero dipole moment (symmetrical).

Structures of $[Pt(NH_3)_2Cl_2]$

This complex (Diamminedichloridoplatinum(II)) exists in two forms:

1. Cis-platin (Cis-isomer)

The two Cl atoms are on the same side. This is a well-known anti-cancer drug.

2. Trans-platin (Trans-isomer)

The two Cl atoms are diagonally opposite to each other.

23

Nomenclature and General Principles

1. Introduction and Catenation

Organic Chemistry: Study of carbon compounds.

Catenation: The unique ability of carbon atoms to bond with each other to form long chains or rings.

2. Classification of Organic Compounds

Acyclic/Open Chain: Straight or branched chains.

Cyclic/Closed Chain: Alicyclic (like cyclohexane) or Aromatic (like Benzene).

3. IUPAC Nomenclature System

The name of an organic compound consists of three parts:

1. Word Root: Number of carbon atoms in the parent chain **(Meth-, Eth-, Prop-, But-, etc.).**

2. Suffix:

Primary: Indicates saturation **(-ane** for **C - C**, -ene for **C = C**, -yne for **C ≡ C).**

Secondary: Indicates functional groups (-ol for alcohol, -al for aldehyde, -oic acid for carboxylic acid).

3. Prefix: Indicates substituent groups like Methyl **($-CH_3$)** or Halogens **(-Cl, -Br).**

4. Electronic Displacement Effects

Inductive Effect (I-effect): Permanent displacement of sigma electrons due to electronegativity difference.

+I groups: **$-CH_3$, $-C_2H_5$** (Electron donating).

-I groups: **$-NO_2$, -CN, -Cl** (Electron withdrawing).

Electromeric Effect (E-effect): Temporary displacement of **pi-**electrons in the presence of an attacking reagent.

Resonance (Mesomeric) Effect: Delocalization of pi-electrons in conjugated systems like Benzene.

5. Fission of a Covalent Bond

Homolytic Fission: Each atom takes one electron, forming **Free Radicals ($A^\bullet$, $B^\bullet$)**

Heterolytic Fission: One atom takes both electrons, forming **Carbocations (C^+)** or **Carbanions (C^-)**.

6. Attacking Reagents

Electrophiles (E^+): Electron-deficient species **(e.g., H^+, $AlCl_3$, BF_3).**

Nucleophiles (Nu^-): Electron-rich species **(e.g., OH^-, NH_3, CN^-).**

7. Types of Organic Reactions

Substitution: One atom is replaced by another **($CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$).**

Addition: Atoms added to double/triple bonds **($CH_2 = CH_2 + H_2 \rightarrow CH_3 - CH_3$).**

Elimination: Atoms removed to form multiple bonds.

Rearrangement: Internal migration of atoms.

8. Isomerism

Structural: Chain, Position, Functional, and Metamerism.

Stereoisomerism: Geometrical **(Cis-Trans)** and Optical.

Top 10 Important Questions

Q1. Explain Catenation in carbon.

Ans. Definition: Catenation is the unique ability of an element to form stable covalent bonds with other atoms of the same element, resulting in long chains, branched chains, or rings.

Carbon shows the maximum catenation power in the periodic table due to two main reasons:

1. Small Size: The small size of carbon allows the nuclei to hold the shared pair of electrons strongly.
2. High Bond Enthalpy: The C-C bond is very strong and stable (348 kJ/mol), which prevents the chains from breaking easily.

Q2. Write the IUPAC name of $CH_3 - CH_2 - CH_2 - OH$

Ans. Propan-1-ol

Naming Breakdown:

- Word Root: 3 Carbon atoms → Prop
- Saturation: Single bonds → an

- Functional Group: Alcohol (-OH) → ol (at position 1)

Q3. Define Nucleophiles with two examples.

Ans. **Definition:** A nucleophile is a chemical species (ion or molecule) that is electron-rich and donates an electron pair to an electron-deficient center (like a nucleus or a positive charge) to form a chemical bond.

- Literal meaning: "Nucleus-loving" (Nucleo = nucleus, phile = loving).

Two Examples

1. **Neutral Nucleophiles:** These have lone pairs of electrons.
   - **Example:** Ammonia ($:NH_3$) or Water ($H_2O:$)
2. **Anionic Nucleophiles:** These carry a negative charge.
   - **Example:** Hydroxide ion (OH^-) or Cyanide ion (CN^-)

Q4. Difference between Inductive and Electromeric effects.

Ans.

Feature	Inductive Effect (I)	Electromeric Effect (E)
Nature	Permanent effect.	Temporary effect (only in presence of a reagent).
Electrons	Displacement of σ (sigma) electrons.	Complete transfer of π (pi) electrons.
Bond Type	Occurs in saturated chains (single bonds).	Occurs in unsaturated systems (double/triple bonds).
Charges	Development of partial charges (δ+, δ-).	Development of complete charges (+, -).
Reagent	No attacking reagent required.	Attacking reagent is mandatory.

Q5. What are Free Radicals? How are they formed?

Ans. **Definition:**

A free radical is an atom or group of atoms that has at least one unpaired electron. Because of this unpaired electron, they are highly reactive and unstable.

- Notation: Represented by a dot (·) over the symbol, e.g., Cl·, $CH_3^{\cdot}$

Process: In a covalent bond (A-B), the shared pair of electrons is split equally. Each atom takes one electron from the bond.

Conditions: This usually happens in the presence of:

- UV Light or Sunlight
- High Temperature (Heat)
- Peroxides (as initiators)

Q6. Explain Geometrical Isomerism with But-2-ene.

Ans. Definition: Geometrical isomerism is a type of stereoisomerism that occurs due to restricted rotation around a carbon-carbon double bond (C=C).

The Two Forms of But-2-ene

In But-2-ene (CH_3-CH=CH-CH_3), the two methyl (-CH_3) groups can be arranged in two different ways:

1. Cis-But-2-ene:
   - The two identical groups (-CH_3) are on the same side of the double bond.
2. Trans-But-2-ene:
   - The two identical groups (-CH_3) are on the opposite sides (diagonally) of the double bond.

Q7. Write one example of an Addition Reaction.

Ans. An addition reaction occurs when two or more molecules combine to form a single larger molecule. This typically happens in unsaturated compounds (containing double or triple bonds).

Example: Hydrogenation of Ethene

In this reaction, Hydrogen (H_2) is added to Ethene (CH_2=CH_2) in the presence of a catalyst like Nickel (Ni) or Platinum (Pt) to form Ethane (CH_3-CH_3).

Chemical Equation

$$CH_2 = CH_2 + H_2 \xrightarrow{Ni/Pt} CH_3 - CH_2$$

- Reactants: Ethene (Unsaturated) + Hydrogen
- Product: Ethane (Saturated)

Q8. Define Functional Group.

Ans.

Functional Group	Formula	Name of Class	Example
Hydroxyl	-OH	Alcohols	Methanol (CH_3OH)
Aldehydic	-CHO	Aldehydes	Ethanal (CH_3CHO)
Carboxyl	-COOH	Carboxylic acids	Ethanoic acid (CH_3COOH)
Ketonic	>C=O	Ketones	Propanone (CH_3COCH_3)

Q9. Explain Resonance in Benzene with a diagram.

Ans. Definition: Resonance is the phenomenon where a single Lewis structure cannot explain all the properties of a molecule. Instead, the molecule is represented by a hybrid of two or more structures.

Key Points in Benzene (C_6H_6):

1. Delocalization: In Benzene, the six π (pi) electrons are not fixed between specific carbon atoms. They are delocalized (spread out) over all six carbon atoms in the ring.
2. Kekulé Structures: Benzene is represented by two structures with alternating double bonds.
3. Resonance Hybrid: The actual structure is a "hybrid" of these two, often drawn as a hexagon with a circle inside.
4. Bond Length: Because of resonance, all C-C bond lengths in benzene are equal (139 pm), which is between a single bond (154 pm) and a double bond 134pm).

Q10. Arrange Carbocations in increasing order of stability (1°, 2°, 3°).

Ans. Increasing Order of Stability

$Methyl < 1^\circ$ **(Primary)** $< 2^\circ$ **(Secondary)** $< 3^\circ$ **(Tertiary)**

Reason for Stability

1. Inductive Effect (+I effect): Alkyl groups (like $-CH_3$) are electron-donating. They push electrons toward the positive carbon, reducing its charge and making it more stable.
2. Hyperconjugation: The more alkyl groups there are, the more C-H sigma bonds can overlap with the empty p-orbital of the carbocation, spreading out the charge.

24 Hydrocarbons

1. Alkanes (Saturated Hydrocarbons)

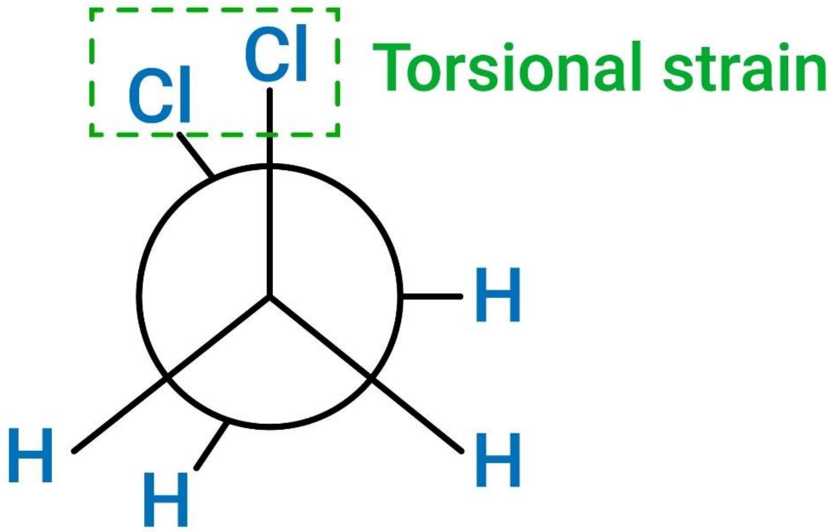
Methods of Preparation:

1. **Wurtz Reaction:** $2RX + 2Na \xrightarrow{\text{Dry Ether}} R - R + 2NaX.$
2. **Decarboxylation:** $RCOONa + NaOH \xrightarrow{CaO,\Delta} R - H + Na_2CO_3.$
3. **Physical Properties:** Boiling point increases with molecular mass and decreases with branching.

Chemical Properties:

- **Halogenation:** $CH_4 + Cl_2 \xrightarrow{hv} CH_3Cl + HCl.$

Conformations of Ethane: Staggered is more stable than Eclipsed.



Eclipsed conformation

2. Alkenes (Unsaturated - Double Bond)

Methods of Preparation:

1. **Dehydrohalogenation:** Alkyl halide + alcoholic KOH → Alkene.
2. **Dehydration of Alcohols:** $C_2H_5OH \xrightarrow{H_2SO_4, 443K} CH_2 = CH_2 + H_2O.$

Physical Properties: Trans-isomers have higher melting points than Cis-isomers.

Chemical Properties:

- **Markovnikov's Rule:** Negative part of reagent adds to C with fewer H atoms.
- **Ozonolysis:** Cleavage of C=C to give Aldehydes/Ketones.

3. Alkynes (Unsaturated - Triple Bond)

Methods of Preparation:

1. **From Calcium Carbide:** $CaC_2 + {}_2H_2O \rightarrow C_2H_2 + Ca(OH)_2$
2. **From Vicinal Dihalides:** Reaction with alcoholic KOH followed by $NaNH_2$

Physical Properties: They are non-polar and insoluble in water.

Chemical Properties:

- **Acidic Nature:** Terminal alkynes react with Sodium metal to release H_2 gas.

$$H-C\equiv C-H+2Na\rightarrow Na-C\equiv C-Na+H_2$$

- **Hydration:** Addition of water in presence of Hg^{2+}/H^+ gives Carbonyl compounds.

4. Aromatic Hydrocarbons (Benzene)

Structure: Cyclic, planar, and shows resonance. It has 6 delocalized π electrons (Huckel's Rule: 4n+2)

Methods of Preparation:

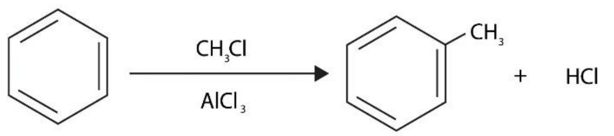
1. **Decarboxylation of Sodium Benzoate:**

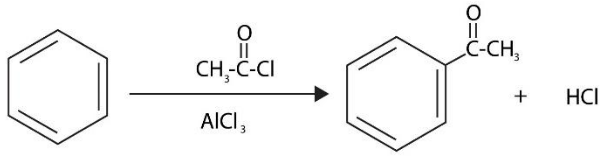
$$C_6H_5COONa+NaOH\xrightarrow{CaO}C_6H_6+\ Na_2CO_3.$$

2. **Reduction of Phenol:** $C_6H_5OH+Zn\xrightarrow{\Delta}C_6H_6+ZnO.$

Chemical Properties (Electrophilic Substitution):

1. **Nitration:** Reaction with conc. HNO_3 + H_2SO_4 to give Nitrobenzene.
2. **Friedel-Crafts Alkylation:** Reaction with R-Cl in presence of anhydrous $AlCl_3$





Top 10 Important Questions (With Answers)

Q1. Name the product formed when Ethyne is passed through a red hot iron tube.

Ans. Benzene (C_6H_6)

Q2. What is the H-C-H bond angle in Methane?

Ans. 109.5°

Q3. Give the reagent for Decarboxylation.

Ans. Soda lime (NaOH + CaO).

Q4. Is Benzene an electrophile or a nucleophile?

Ans. Nucleophile (due to π electron cloud).

Q5. Write the IUPAC name of the simplest Alkyne?

Ans. Ethyne

Q6. Explain the acidic nature of Ethyne with a reaction.

Ans. $H - C \equiv C - H + 2Na \rightarrow Na - C \equiv C - Na + H_2$

Q7. How is Benzene prepared from Phenol?

Ans. By heating Phenol with Zinc dust. $C_6H_5OH + Zn \rightarrow C_6H_6 + ZnO$

Q8. State Markovnikov's rule with an example.

Ans. $CH_3 - CH = CH_2 + HBr \rightarrow CH_3 - CH(Br) - CH_3$. Br^- goes to carbon with fewer H.

Q9. Explain Nitration of Benzene.

Ans. Benzene reacts with conc. HNO_3 and H_2SO_4 at 333K to form Nitrobenzene.

Q10. What is Ozonolysis of Ethene? Write the reaction.

Ans. $CH_2 = CH_2 + O_3 \rightarrow \text{Ozonide} \xrightarrow{Zn/H_2O} 2HCHO$ (Formaldehyde).

Compounds of Carbon Containing Halogens

1. Introduction

Haloalkanes (alkyl halides) and Haloarenes (aryl halides) are formed by replacing hydrogen atoms of hydrocarbons with halogen atoms (F, Cl, Br, I).

2. Methods of Preparation (Reactions)

From Alcohols: Reaction with HX, PCl_3, PCl_5, or $SOCl_2$

R - OH + $SOCl_2$ → R - Cl + SO_2 + HCl

(Best method as byproducts are gases).

From Hydrocarbons:

- **Free Radical Halogenation:** $CH_4 + Cl_2 \xrightarrow{hv} CH_3Cl + HCl.$
- **Electrophilic Substitution (for Haloarenes):** Benzene $Cl_2 \xrightarrow{FeCl_3}$ Chlorobenzene

Sandmeyer's Reaction: Preparation of haloarenes from Benzene diazonium salt using Cu_2Cl_2 or Cu_2Br_2

Finkelstein Reaction (Halogen Exchange): $R - X + NaI \xrightarrow{\text{Acetone}} R - I + NaX.$

3. Physical Properties

Solubility: Although polar, they are insoluble in water because they cannot form hydrogen bonds with water molecules.

Boiling Point: Increases with atomic mass of halogen (R - I > R - Br > R - Cl > R - F). For isomeric halides, boiling point decreases with branching.

4. Chemical Properties (Reactions)

Nucleophilic Substitution (S_N1 and S_N2):

- **S_N2:** Single step, inversion of configuration, order: $1^\circ > 2^\circ > 3^\circ$.
- **S_N1:** Two steps, formation of carbocation, order: $3^\circ > 2^\circ > 1^\circ$.

Elimination (Dehydrohalogenation): Alkyl halide + alc. KOH → Alkene (Follows Saytzeff's Rule).

Wurtz-Fittig Reaction: $C_6H_5Br + 2Na + CH_3Br \xrightarrow{\text{Ether}} C_6H_5 - CH_3$ (Toluene).

Top 10 Important Questions (With Answers)

Q1. Why are haloalkanes insoluble in water?

Ans: Because they cannot form hydrogen bonds with water.

Q2. What is the order of reactivity of alkyl halides towards S_N2 reaction?

Ans. $1° > 2° > 3°$.

Q3. Give the IUPAC name of CH_3 - CH(Cl) - CH_3

Ans. 2-chloropropane

Q4. Which reagent is used in Finkelstein reaction?

Ans: NaI in dry acetone.

Q5. Name the major product of dehydrohalogenation of 2-bromobutane.

Ans. But-2-ene

Q6. Explain Sandmeyer's reaction with a balanced equation.

Ans. Benzene diazonium chloride reacts with Cu_2Cl_2/HCl to give Chlorobenzene.

$$C_6H_5N_2Cl \xrightarrow{Cu_2Cl_2/HCl} C_6H_5Cl + N_2$$

Q7. Distinguish between S_N1 and S_N2 mechanisms.

Ans. S_N1 is a two-step process involving a carbocation, while S_N2 is a single-step process with a transition state.

Q8. Write the Wurtz-Fittig reaction for the preparation of Toluene.

Ans. $C_6H_5Br + 2Na + CH_3Br \xrightarrow{\text{Ether}} C_6H_5 - CH_3 + 2NaBr.$

Q9. What is Saytzeff's Rule? Give an example.

Ans. In elimination reactions, the more substituted alkene is the major product.

$$CH_3 - CH_2 - CH(Br) - CH_3 \xrightarrow{\text{alc. } KOH} CH_3 - CH = CH - CH_3$$

Q10. How will you prepare Iodoethane from Bromoethane?

Ans. By Finkelstein reaction: $C_2H_5Br + NaI \xrightarrow{\text{Acetone}} C_2H_5I + NaBr.$

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Alcohols, Phenols and Ethers

1. Alcohols (R - OH)

A. Methods of Preparation (Reactions)

From Alkenes (Acid Catalyzed Hydration):

$$CH_2 = CH_2 + H_2O \xrightarrow{H^+} CH_3 - CH_2 - OH$$

Reduction of Carbonyl Compounds: Aldehydes give 1° alcohols and Ketones give 2° alcohols using $LiAlH_4$ or $NaBH_4$

$$R - CHO \xrightarrow{LiAlH_4} R - CH_2OH$$

From Grignard Reagents: Reaction of R-MgX with aldehydes/ketones followed by hydrolysis.

$$HCHO + RMgX \rightarrow RCH_2OMgX \xrightarrow{H_2O} RCH_2OH \ (1^\circ \text{ Alcohol})$$

B. Physical Properties

Boiling Point: Increases with molecular mass. Alcohols have higher boiling points than hydrocarbons of comparable mass due to Intermolecular Hydrogen Bonding.

Solubility: Lower alcohols are highly soluble in water due to their ability to form hydrogen bonds with water molecules.

C. Chemical Properties (Reactions)

Dehydration: Heating with conc. H_2SO_4 at 443K produces Alkenes.

$$CH_3CH_2OH \xrightarrow{H_2SO_4, 443K} CH_2 = CH_2 + H_2O$$

Oxidation:

- 1° Alcohol $\xrightarrow{PCC}$ Aldehyde.
- 1° Alcohol $\xrightarrow{KMnO_4}$ Carboxylic Acid.

Lucas Test (Distinction):

- 3° Alcohol: Immediate turbidity.
- 2° Alcohol: Turbidity after 5 minutes.
- 1° Alcohol: No turbidity at room temperature.

2. Phenols (Ar - OH)

A. Methods of Preparation

1. **From Benzene Diazonium Salt:** Hydrolysis with warm water.

$$C_6H_5N_2Cl + H_2O \xrightarrow{\text{Warm}} C_6H_5OH + N_2 + HCl$$

2. **From Haloarenes (Dow's Process):** Chlorobenzene + NaOH at 623K and 300 atm.
3. **From Cumene:** Oxidation of cumene to cumene hydroperoxide followed by acid hydrolysis yields Phenol and Acetone**.**

B. Physical Properties

Phenols are colorless liquids or crystalline solids.

They turn reddish due to slow atmospheric oxidation.

High boiling points due to strong intermolecular hydrogen bonding.

C. Chemical Properties

1. **Acidity:** Phenols are more acidic than alcohols because the Phenoxide ion is resonance stabilized.
2. **Reimer-Tiemann Reaction:**

$$C_6H_5OH + CHCl_3 + aq.NaOH \rightarrow \text{Salicylaldehyde}$$

3. **Kolbe's Reaction:**

$$C_6H_5OH + NaOH + CO_2 \rightarrow \text{Salicylic acid}$$

3. Ethers (R – O - R')

A. Methods of Preparation

Williamson Synthesis: Reaction of alkyl halide with sodium alkoxide.

$$R - X + R' - ONa \rightarrow R - O - R' + NaX$$

Dehydration of Alcohols:

$$2CH_3CH_2OH \xrightarrow{H_2SO_4, 413K} C_2H_5 - O - C_2H_5.$$

B. Physical & Chemical Properties

- **Ethers have much lower boiling points than alcohols (no H-bonding).**
- **Cleavage by HI:** $R - O - R' + HI \rightarrow R - OH + R' - I.$

Top 10 Important Questions (With Answers)

Q1. Which alcohol gives immediate turbidity with Lucas reagent?

Ans. Tertiary (3°) alcohol.

Q2. Name the product formed in Kolbe's reaction.

Ans. Salicylic acid

Q3. Why is Phenol more acidic than Ethanol?

Ans. Due to resonance stabilization of phenoxide ion.

Q4. Name the reagent used to convert 1° alcohol to aldehyde.

Ans. PCC (Pyridinium chlorochromate)

Q5. What is the IUPAC name of Phenol?

Ans. Benzenol

Q6. Explain Williamson synthesis with an equation

Ans. $CH_3Br + C_2H_5Ona \rightarrow CH_3 - O - C_2H_5 + NaBr$

Q7. Write the chemical equation for Reimer-Tiemann reaction.

Ans. $C_6H_5OH + CHCL_3 + 3NaOH \rightarrow$ Salicylaldehyde $+ 3NaCl + 2H_2O$

Q8. How will you distinguish between Ethanol and Propan-2-ol?

Ans. By Lucas test: Propan-2-ol (2°) gives turbidity in 5 mins, Ethanol (1°) does not.

Q9. Explain the preparation of Phenol from Cumene.

Ans. Cumene $\xrightarrow{O_2}$ Cumene hydroperoxide $\xrightarrow{H^+/H_2O}$ Phenol + Acetone.

Q10. Write the mechanism of dehydration of Ethanol to Ethene.

Ans. Protonation → Carbocation formation → Elimination of proton.

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Aldehydes, Ketones and Carboxylic Acids

1. Aldehydes and Ketones (Carbonyl Compounds)

A. Methods of Preparation (Reactions)

Oxidation of Alcohols: 1° alcohols give aldehydes and 2° alcohols give ketones using $K_2Cr_2O_7$ or PCC.

$$R - CH_2OH \xrightarrow{PCC} R - CHO$$

Dehydrogenation of Alcohols: Passing alcohol vapors over heated Copper at 573K.

$$CH_3CH_2OH \xrightarrow{Cu,573K} CH_3CHO + H_2$$

Hydration of Alkynes: Addition of water to ethyne in presence of Hg^{2+}/H^+

$$CH \equiv CH + H_2O \xrightarrow{HgSO_4/H_2SO_4} CH_3CHO$$

Rosenmund Reduction: Hydrogenation of acid chlorides using $Pd/BaSO_4$

$$R - COCl + H_2 \xrightarrow{Pd/BaSO_4} R - CHO + HCl$$

B. Physical Properties

Boiling Point: Higher than hydrocarbons but lower than alcohols of comparable mass because they cannot form intermolecular hydrogen bonds.

Solubility: Lower members (Formaldehyde, Acetaldehyde) are soluble in water due to hydrogen bonding with water molecules.

C. Chemical Properties (Reactions)

Nucleophilic Addition: Addition of HCN, $NaHSO_3$, and Grignard reagents (RMgX).

Reduction:

- **Clemmensen Reduction:** Using Zn - Hg/HCl to form alkanes.
- **Wolff-Kishner Reduction:** Using NH_2NH_2/KOH

Aldol Condensation: Occurs in aldehydes/ketones with α-hydrogen in presence of dilute alkali.

Cannizzaro Reaction: Occurs in aldehydes with **no** α-hydrogen (like Formaldehyde).

$$2HCHO + \text{conc. } NaOH \rightarrow CH_3OH + HCOONa$$

2. Carboxylic Acids (R-COOH)

A. Methods of Preparation

1. **Oxidation:** Primary alcohols or aldehydes with $KMnO_4$
2. **Hydrolysis of Nitriles:** $R - CN \xrightarrow{H_3O^+} R - COOH.$

B. Physical & Chemical Properties

Physical: Highest boiling points among organic compounds due to extensive **dimer formation** via hydrogen bonding.

Acidity: More acidic than phenols. Electron withdrawing groups (like -Cl, $-NO_2$) increase acidity.

Esterification: Reaction with alcohols in presence of H^+ to form Esters (RCOOR').

Top 10 Important Questions (With Answers)

Q1. Name the product of Rosenmund reduction of Acetyl chloride.

Ans. Acetaldehyde

Q2. Which test distinguishes Aldehydes from Ketones?

Ans. Tollen's test or Fehling's test.

Q3. Why do carboxylic acids have higher boiling points than alcohols?

Ans. Due to the formation of stable hydrogen-bonded dimers.

Q4. Give one example of an aldehyde that does not undergo Aldol condensation.

Ans. Formaldehyde (HCHO) or Benzaldehyde.

Q5. What is the IUPAC name of CH_3COCH_3?

Ans. Propanone

Q6. Explain Clemmensen reduction with an equation

Ans. $CH_3CHO + 4[H] \xrightarrow{Zn-Hg/HCl} CH_3 - CH_3 + H_2O.$

Q7. Write the chemical equation for Cannizzaro reaction.

Ans. $2HCHO + NaOH \rightarrow CH_3OH + HCOONa$

Q8. How will you prepare Acetaldehyde from Ethyne?

Ans. $CH \equiv CH + H_2O \xrightarrow{Hg^{2+}/H^+} CH_3CHO.$

Q9. What is Aldol Condensation? Give reaction for Ethanal.

Ans. $2CH_3CHO \xrightarrow{\text{dil. } NaOH} CH_3 - CH(OH) - CH_2 - CHO$ (Aldol)

Q10. Explain Esterification with a balanced equation

$CH_3COOH + C_2H_5OH \xrightarrow{H^+} CH_3COOC_2H_5 + H_2O.$

COMPOUNDS OF CARBON CONTAINING NITROGEN

1. Amines (R-NH_2)

Amines are derivatives of ammonia (NH_3) where one or more hydrogen atoms are replaced by alkyl or aryl groups.

A. Methods of Preparation (Reactions)

Reduction of Nitro Compounds: Nitrobenzene reduces to Aniline using Sn/HCl or Fe/HCl.

$$C_6H_5NO_2 + 6[H] \xrightarrow{Sn/HCl} C_6H_5NH_2 + 2H_2O$$

Hoffmann Bromamide Degradation: Reaction of an amide with Br_2 and NaOH to give an amine with one less carbon.

$$R - CONH_2 + Br_2 + 4NaOH \rightarrow R - NH_2 + Na_2CO_3 + 2NaBr + 2H_2O$$

Gabriel Phthalimide Synthesis: Used specifically for preparing **primary aliphatic amines**.

Reduction of Nitriles: $R - CN \xrightarrow{LiAlH_4} R - CH_2NH_2$.

B. Physical Properties

Boiling Point: Primary and secondary amines have higher boiling points than alkanes due to **intermolecular hydrogen bonding**. Order: 1° > 2° > 3° (tertiary has no H-bond).

Basicity: Amines are basic due to the lone pair on Nitrogen. Aliphatic amines are more basic than Ammonia, but **Aniline is less basic** than Ammonia because the lone pair is delocalized in the benzene ring.

C. Chemical Properties (Reactions)

Carbylamine Reaction: Primary amines react with $CHCl_3$ and alcoholic KOH to give foul-smelling isocyanides.

$R\text{-}NH_2 + CHCl_3 + 3KOH \rightarrow R\text{-}NC + 3KCl + 3H_2O$

Reaction with Nitrous Acid (HNO_2):

- Aliphatic amines give alcohols and N_2 gas.
- Aniline gives **Benzene Diazonium Chloride** at 273-278K (Diazotization).

2. Nitro Compounds (R-NO_2)

A. Preparation & Properties

Nitration of Benzene: Reaction with conc. HNO_3 + H_2SO_4

Reduction: Nitro groups can be reduced to amines or hydroxylamines depending on the medium.

Top 10 Important Questions (With Answers)

Q1. Which test is used to identify primary amines?

Ans. Carbylamine test

Q2. Why is Aniline less basic than Methylamine?

Ans. Due to delocalization of lone pair in the benzene ring.

Q3. What is the product of Hoffmann Bromamide reaction of Benzamide?

Ans. Aniline

Q4. Name the reagent used for Diazotization.

Ans. $NaNO_2$ + HCl at 273-278K.

Q5. Write the IUPAC name of CH_3-NH-CH_2CH_3

Ans. N-Methylethanamine

Q6. Explain Hoffmann Bromamide Degradation with an equation.

Ans. $CH_3CONH_2 + Br_2 + 4NaOH \rightarrow CH_3NH_2 + Na_2CO_3 + 2NaBr + 2H_2O$

Q7. Write the chemical reaction for Carbylamine test

Ans. $C_2H_5NH_2 + CHCl_3 + 3KOH \rightarrow C_2H_5NC + 3KCl + 3H_2O$

Q8. How will you prepare Aniline from Nitrobenzene?

Ans. By reduction using Sn/HCl. $C_6H_5NO_2 + 6[H] \rightarrow C_6H_5NH_2 + 2H_2O$

Q9. What is Gabriel Phthalimide Synthesis? Why is it preferred for 1° amines?

Ans. It avoids the formation of 2° and 3° amines, giving pure 1° aliphatic amine.

Q10. Explain the Diazotization reaction of Aniline.

Ans. Aniline reacts with HNO_2 (prepared from $NaNO_2$/HCl) at 0-5°C to give Benzene Diazonium Chloride.

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Biomolecules

1. Carbohydrates

Carbohydrates are optically active polyhydroxy aldehydes or ketones. They are the primary source of energy for living organisms.

Classification:

- **Monosaccharides:** Simplest sugars that cannot be further hydrolyzed (e.g., Glucose, Fructose).
- **Oligosaccharides:** Yield 2–10 monosaccharide units on hydrolysis (e.g., Sucrose, Maltose).
- **Polysaccharides:** High molecular mass polymers yielding many monosaccharide units (e.g., Starch, Cellulose, Glycogen).

Glucose ($C_6H_{12}O_6$): An aldohexose. It exists in an open-chain form and cyclic forms (α and β anomers).

2. Proteins and Amino Acids

Proteins are polymers of α-**amino acids** joined by **peptide bonds** (-CONH-).

Zwitterion: In aqueous solution, the carboxyl group loses a proton and the amino group accepts one, forming a dipolar ion called a Zwitterion.

Structure of Proteins:

1. **Primary:** Linear sequence of amino acids.
2. **Secondary:** Folding into α-helix or β-pleated sheets due to hydrogen bonding.
3. **Tertiary:** Overall 3D shape of the polypeptide chain.
4. **Quaternary:** Spatial arrangement of two or more polypeptide chains.

Denaturation: Loss of biological activity and 3D structure due to changes in temperature or pH.

3. Nucleic Acids

Nucleic acids are polymers of nucleotides that store and transmit genetic information.

DNA (Deoxyribonucleic Acid): Contains deoxyribose sugar and bases: Adenine (A), Guanine (G), Cytosine (C), and **Thymine (T)**. It has a **Double Helix** structure.

RNA (Ribonucleic Acid): Contains ribose sugar and bases: A, G, C, and **Uracil (U)**. It is usually single-stranded.

4. Vitamins and Hormones

Vitamins: Organic compounds required in small amounts for health.

- **Water-soluble:** B-complex and C.
- **Fat-soluble:** A, D, E, and K.

Hormones: Chemical messengers produced by endocrine glands (e.g., Insulin, Adrenaline).

Top 10 Important Questions (With Answers)

Q1. What is the sugar present in RNA?

Ans. Ribose

Q2. Which linkage joins amino acids in a protein?

Ans. Peptide bond

Q3. Name a fat-soluble vitamin.

Ans. Vitamin A (or D, E, K).

Q4. What is the simplest monosaccharide?

Ans. Glucose

Q5. Name the base found in DNA but not in RNA

Ans. Thymine

Q6. Define Zwitterion with its structure.

Ans. It is a dipolar ion formed when the -COOH group loses a proton to the $-NH_2$ group within an amino acid. Structure: H_3N^+-CHR-COO^-

Q7. Explain the difference between DNA and RNA

Ans. DNA contains deoxyribose sugar and thymine, and is double-stranded. RNA contains ribose sugar and uracil, and is single-stranded.

Q8. What is the denaturation of proteins? Give one example

Ans. Loss of a protein's 3D structure and biological activity due to heat or chemicals. Example: Coagulation of egg white on boiling.

Q9. Classify carbohydrates based on their hydrolysis behavior.

Ans. Monosaccharides (no hydrolysis), Oligosaccharides (2-10 units), and Polysaccharides (many units).

Q10. What are Nucleotides and Nucleosides?

Ans. Nucleoside = Sugar + Base; Nucleotide = Sugar + Base + Phosphate group.

SOAP, DETERGENTS AND POLYMERS

1. Cleansing Agents: Soaps and Detergents

Soaps: These are sodium or potassium salts of long-chain fatty acids (e.g., stearic acid). They are made by **Saponification**, which is the alkaline hydrolysis of oils or fats.

Synthetic Detergents: These are sodium salts of long-chain alkyl hydrogen sulphates or alkyl benzene sulphonic acids. Unlike soaps, they work effectively even in **hard water** because their calcium and magnesium salts are water-soluble.

Cleansing Mechanism: Both contain a **hydrophilic (water-loving) head** and a **hydrophobic (water-fearing) tail**. In water, they form **micelles**, where the tails trap grease/dirt and the heads stay in contact with water.

2. Introduction to Polymers

Definitions:

- **Polymer:** A giant molecule formed by linking many small units called monomers.
- **Homopolymer:** Made from only one type of monomer (e.g., Polyethene).
- **Copolymer:** Made from different types of monomers (e.g., Buna-S).

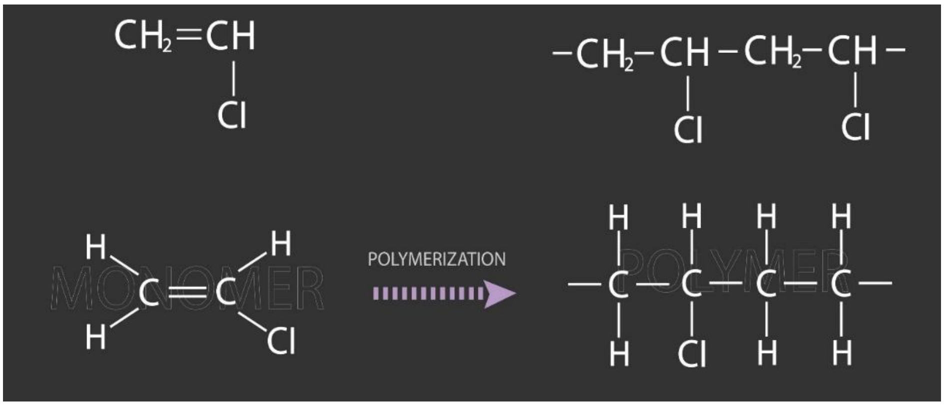
3. Polymerization Reactions

A. Addition Polymerization (Chain Growth)

This involves the repeated addition of monomers containing double or triple bonds without losing any molecules.

Mechanism: It usually proceeds via a **Free Radical Mechanism** involving three steps: Initiation, Propagation, and Termination.

Example: Polyethene $n(CH_2 = CH_2) \xrightarrow{\text{Polymerization}} -[CH_2 - CH_2]_n-$



B. Condensation Polymerization (Step Growth)

This involves the linking of monomers with the elimination of small molecules like H_2O or NH_3

- **Example: Nylon-6,6** (from Adipic acid and Hexamethylenediamine).

$$n[H_2N(CH_2)_6NH_2] + n[HOOC(CH_2)_4COOH] \rightarrow -[NH(CH_2)_6NHCO(CH_2)_4CO]_n$$

4. Biodegradable Polymers

Standard synthetic polymers cause pollution as they don't degrade. **Biodegradable polymers** like **PHBV** or **PGA** are designed to break down via enzymatic action.

Top 10 Questions with Answers (Short & Long)

Q1. What is Saponification?

Ans. The process of making soap by alkaline hydrolysis of fats/oils.

Q2. Define Monomer.

Ans. The simple molecules that link to form a polymer.

Q3. Why are detergents better than soaps for hard water?

Ans. Detergents form water-soluble salts with Ca^{2+} and Mg^{2+} ions.

Q4. Give one example of a natural polymer.

Ans. Cellulose or Starch.

Q5. What is PHBV used for?

Ans. In orthopedic devices and controlled drug release.

Q6. Explain the mechanism of Addition Polymerization.

Ans. It involves Initiation (forming free radicals), Propagation (growing the chain), and Termination (combining chains).

Q7. Write the chemical equation for the synthesis of Nylon-6,6

Ans. $n[H_2N(CH_2)_6NH_2] + n[HOOC(CH_2)_4COOH] \xrightarrow{\Delta} \text{Nylon-6,6} + 2nH_2O.$

Q8. Differentiate between Thermoplastics and Thermosetting polymers

Ans. Thermoplastics can be remoulded on heating; Thermosetting polymers set permanently after one heating.

Q9. What is Vulcanization of rubber?

Ans. Heating raw rubber with sulphur to create cross-links, improving strength and elasticity.

Q10. Explain the cleansing action of soap micelles.

Ans. The hydrophobic tail traps grease, while the hydrophilic head stays in water. Stirring breaks grease into droplets forming an emulsion that washes away.