CHEMISTRY Mastery for NET

Complete Chapter-wise Guide

Chapters (as per PTB & NET syllabus):

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Chapter 1: Basic Concepts of Chemistry

Key Topics Covered:

- Atomic and Molecular Masses
- Molar Mass and Mole
- Avogadro's Number
- Empirical and Molecular Formulas
- Limiting Reactant and Excess Reactant
- Percentage Composition
- Stoichiometry and Yield

Core Definitions and Ideas:

🔷 Atomic Mass

- Mass of a single atom, relative to 1/12th of C-12.
- **Unit:** atomic mass unit (amu or u)
- $1 \text{ u} = 1.66 \times 10^{-27}$ kg

🔷 Molecular Mass

- Sum of atomic masses of all atoms in a molecule.
  👉 Example: $H_2O = 2\times1 + 16 =$ **18 u**

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🔷 Formula Mass

- Used for ionic compounds (e.g., $NaCl = 23 + 35.5 = 58.5$ u)

🔷 Molar Mass

- Mass of one mole (6.022×10^{23} particles) of a substance in grams
  👉 H_2O molar mass = 18 g/mol

🔢 Avogadro's Number

- **$N_a = 6.022 \times 10^{23}\ mol^{-1}$**
- No. of atoms/molecules/ions/electrons in one mole

🔷 The Mole Concept

Number of Moles (n):
n = Given Mass ÷ Molar Mass

No. of Particles:
$= n \times N_a$

Empirical vs Molecular Formula:

Term	Description
Empirical	Simplest whole-number ratio of elements
Molecular	Actual number of atoms of each element

To find molecular formula:
Molecular Formula = Empirical Formula × (Molar Mass ÷ Empirical Mass)

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Percentage Composition:

% Composition = (Mass of Element ÷ Molar Mass of Compound) × 100

🧮 **Example:**
Find % of C in CO_2
→ M(C) = 12, M(O_2) = 32 → M(CO_2) = 44
→ %C = (12 ÷ 44) × 100 = **27.27%**

Limiting Reactant and Yield:

Steps to Determine Limiting Reactant:

1. Convert grams to moles
2. Use mole ratio from balanced equation
3. Compare required vs given moles

Types of Yield:

- **Theoretical Yield:** Predicted from stoichiometry
- **Actual Yield:** Measured experimentally
- **% Yield = (Actual ÷ Theoretical) × 100**

Gas Volume Calculations (At STP):

- 1 mole of **any gas** at STP = **22.4 dm^3**
  👉 **Volume = n × 22.4**

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Key Formulas:

Concept	**Formula**
Moles	**n = mass ÷ molar mass**
Atoms/molecules	**$N = n \times N_a$**
Mass	**mass = n × molar mass**
Volume at STP	**$V = n \times 22.4\ dm^3$**
Empirical formula ratio	**Divide % by atomic mass, simplify**
% Composition	**(mass of element ÷ molar mass) × 100**
% Yield	**(actual ÷ theoretical) × 100**

Top 10 NET MCQ Tips:

1. Always **convert grams to moles** before using mole ratio
2. Be careful with units: g ↔ mol ↔ dm^3 ↔ cm^3
3. **Empirical formula problems** require simplification of smallest ratios
4. Avogadro's number = particles/mole – memorize it
5. Check if question is asking **atoms or molecules**
6. For volume problems, check if conditions are **at STP**
7. Use **limiting reactant** for yield questions
8. Practice calculating **% composition** quickly
9. Molar mass = atomic mass × number of atoms
10. Don't forget **diatomic gases** like O_2, H_2, N_2 (important for mass and mole calculations)

Past NET MCQs Examples:

Q1:
How many atoms are there in 2 moles of Na?
A. 1.2×10^{23} **B.** 6.02×10^{23} **C.** 1.2×10^{24} **D.** 2×10^{23}
✅ **Correct: C**

Q2:
The empirical formula of a compound is CH_2O and its molar mass is 180 g/mol. What is its molecular formula?
A. CH_2O **B.** $C_2H_4O_2$ **C.** $C_6H_{12}O_6$ **D.** $C_3H_6O_3$
✅ **Correct: C**

Q3:
What is the molar mass of H_2SO_4?
A. 96 g/mol **B.** 98 g/mol **C.** 100 g/mol **D.** 94 g/mol
✅ **Correct: B**

Q4:
Which one is the limiting reactant if 10g of H_2 reacts with 80g of O_2?
$2H_2 + O_2 \rightarrow 2H_2O$
✅ **H_2 is limiting reactant**

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Chapter 2: Atomic Structure

Core Concepts Overview:

This chapter deals with the structure of atoms, discovery of subatomic particles, Bohr's atomic model, quantum numbers, orbitals, and electron configurations – **frequently asked in NET MCQs**.

Key Discoveries in Atomic Structure:

Scientist	**Discovery**
J.J. Thomson	Electron (Cathode Ray Tube)
Goldstein	Proton (Canal Rays)
Chadwick	Neutron
Rutherford	Nucleus (Gold Foil Experiment)
Bohr	Quantized Energy Levels
Schrödinger	Quantum Mechanical Model (orbitals)

Important Equations and Constants

Concept	**Equation**
Angular Momentum (Bohr)	$\mathbf{mvr = nh \div 2\pi}$
Energy of nth orbit	$\mathbf{E_n = -13.6 \times Z^2 \div n^2}$ eV
Radius of nth orbit	$\mathbf{r_n = 0.529 \times n^2 \div Z}$ Å
Speed of electron in orbit	$\mathbf{v_n = 2.18 \times 10^6 \times Z \div n}$ m/s

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Concept	**Equation**
Energy of photon emitted	**$\Delta E = h \times \nu = E\_final - E\_initial$**
Rydberg Equation	**$1 \div \lambda = R \times (1 \div n_1^2 - 1 \div n_2^2)$**
Max electrons per shell	**$2n^2$**
No. of orbitals per shell	**n^2**

Constants:

- **$h = 6.63 \times 10^{-34}$ J·s** (Planck's constant)
- **$R = 1.097 \times 10^{7}$ m^{-1}** (Rydberg constant)
- **$e = 1.6 \times 10^{-19}$ C**
- **1 eV = 1.6×10^{-19} J**

Bohr's Postulates (Hydrogen Atom):

1. Electrons revolve in **fixed energy levels** (quantized orbits)
2. Angular momentum is **quantized**
3. Energy is absorbed or emitted only when **electron transitions between levels**
4. Only certain orbits allowed → explain hydrogen's **line spectrum**

Hydrogen Spectral Series (Rydberg Series):

Series	Final Level (n_1)	Region
Lyman	1	Ultraviolet
Balmer	2	Visible

Series	Final Level (n_1)	Region
Paschen	3	Infrared
Brackett	4	Infrared
Pfund	5	Infrared

Example:
For **Balmer series**, $n_1 = 2$, $n_2 = 3, 4, 5, ...$

Quantum Numbers:

Symbol	Name	Role	Values
n	Principal	Energy level, size	1, 2, 3, ...
l	Azimuthal	Shape of orbital	0 to n–1 (s, p, d, f)
m	Magnetic	Orientation in space	–l to +l
s	Spin	Direction of spin	+½ or –½

Electron Configurations:

Rules:

- **Aufbau Principle** – Fill lowest energy orbital first
- **Pauli Exclusion** – No 2 electrons have same 4 quantum numbers
- **Hund's Rule** – Pairing only occurs after each orbital has one electron

Example:

• O = $1s^2\ 2s^2\ 2p^4$
• Fe = [Ar] $4s^2\ 3d^6$

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Shortcut Order (Energy Levels):
1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s...

Important Conceptual Highlights:

- Electrons are **negatively charged**, occupy orbitals around nucleus
- Proton number = Atomic number (Z)
- Neutrons affect **mass**, not charge
- Orbitals are **regions of probability**, not fixed paths
- **1 orbital** = max **2 electrons**
- Shapes of orbitals:
  • **s = spherical**, **p = dumbbell**, **d = cloverleaf**

Graphs / Visuals to Remember:

1. **Energy level diagram** (Hydrogen atom with decreasing negative energy)
2. **Electron configuration diagram (Aufbau order)**
3. **Shapes of orbitals**

Common NET MCQ Concepts:

1. **Energy levels** are quantized – electron can't be between levels
2. **Lower orbit = more negative energy = more stable**
3. For **hydrogen**, spectral lines arise when **electrons fall back** to lower levels
4. Use $\Delta E = h\nu = hc \div \lambda$ when photon wavelength/frequency is given
5. Understand difference between **orbital** and **orbit**
6. **Pauli principle** prevents more than 2 electrons in one orbital

7. **Hund's rule** explains unpaired electrons in p/d orbitals
8. Transition energy = **E_high – E_low** → positive value (energy emitted)

Past NET MCQs:

Q1:
Energy of an electron in the 3rd orbit of hydrogen is:
A. –13.6 eV **B.** –1.51 eV **C.** –3.4 eV **D.** –0.85 eV
✅ **Correct: D**
$E_n = -13.6 \div 9 = -1.51$ eV

Q2:
How many orbitals are in the n = 3 shell?
A. 3 **B.** 6 **C.** 9 **D.** 18
✅ **Correct: C**
$n^2 = 3^2 = 9$ orbitals

Q3:
Which rule is violated in configuration $1s^2\ 2s^2\ 2p^4$ (↑↓ ↑↓ ↑ ↑)?
A. Pauli
B. Hund
C. Aufbau
D. All
✅ **Correct: B** – Hund's Rule violated (should have all ↑ first)

Q4:
The maximum number of electrons in a d-subshell is:
A. 2 **B.** 6 **C.** 10 **D.** 14
✅ **Correct: C**

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Chapter 3: Chemical Bonding and Structure

Core Concepts:

This chapter is highly important for NET and includes **Ionic and Covalent bonding**, **Lewis structures**, **Hybridization**, **Molecular shapes (VSEPR)**, and **Electronegativity**.

🔷 1. Types of Chemical Bonds

Type	Formed Between	Electron Behavior
Ionic	Metal + Non-metal	Electron Transfer
Covalent	Non-metal + Non-metal	Electron Sharing
Coordinate Covalent	Lone pair donor → empty orbital acceptor	Shared pair from one atom

Ionic Bond:
Formed by **complete transfer** of electrons (e.g., NaCl)

Covalent Bond:
Formed by **sharing** of electrons (e.g., H_2O, Cl_2)

Coordinate Bond (Dative):
Both shared electrons come from **same atom** (e.g., NH_4^+)

🔷 2. Octet Rule

Atoms tend to gain, lose, or share electrons to complete **8 electrons** in outer shell
→ Exceptions: **H (2 e^-), Be (4), B (6)**

🔷 3. Lewis Structures

- Show **bonding** and **lone pairs**
- Count total valence electrons
- Satisfy octet/duet rules

Example: H_2O
Oxygen has 2 bonds with H and 2 lone pairs → bent shape

🔷 4. Bond Characteristics

Property	**Ionic**	**Covalent**
Melting point	High	Low/moderate
Solubility	Soluble in water	Soluble in non-polar solvents
Conductivity	Conducts in molten/solution	Generally poor
Physical state	Solid (crystal)	Gases/liquids/solids

🔷 5. Electronegativity (EN)

- **Ability of atom to attract shared electrons**
- Measured on Pauling scale
- **F > O > N > Cl > Br > I > S > C > H**

Bond Type Prediction (by ΔEN):

ΔEN	**Bond Type**
0 – 0.4	Non-polar covalent
0.5 – 1.7	Polar covalent
> 1.7	Ionic

🔷 6. Bond Polarity & Dipole Moment

- **Polar bonds** have unequal sharing of e^-
- **Dipole moment (μ)**: Indicates molecular polarity
  $\mu = q \times d$

🔷 7. Bond Energy, Length & Order

- **Bond Order:** Number of bonds (1 = single, 2 = double…)
- **Bond Length:** Inversely related to bond order
- **Bond Energy:** Energy to break bond
  Higher bond order → **shorter and stronger bond**

🔷 8. Valence Shell Electron Pair Repulsion (VSEPR) Theory

Predict **shape of molecules** based on electron repulsion.

Electron Pairs	Shape	Example
2	Linear	CO_2
3	Trigonal planar	BF_3
4	Tetrahedral	CH_4
3 bonding + 1 lone pair	Trigonal pyramidal	NH_3
2 bonding + 2 lone pair	Bent	H_2O

🔷 9. Hybridization

Type	Orbitals mixed	Geometry	Example
sp	1 s + 1 p	Linear (180°)	$BeCl_2$
sp^2	1 s + 2 p	Trigonal planar (120°)	BF_3
sp^3	1 s + 3 p	Tetrahedral (109.5°)	CH_4

Important Equations:

- **Dipole Moment (μ):**
  $\mu = q \times d$ (C·m)
  • q = charge, d = bond length
- **Formal Charge:**
  FC = valence – (non-bonding + ½ bonding)

Key NET Tips:

1. ΔEN helps predict bond type: **>1.7 = ionic**
2. Lone pairs cause **bond angle reduction** (VSEPR)
3. **Covalent compounds** don't conduct electricity
4. **Polar molecules** have non-zero dipole moment
5. **Hybridization** is based on number of electron domains
6. CO_2 = linear (2 domains), NH_3 = pyramidal (3 bonds + 1 lone)
7. Lewis structure: check **total valence e^-**
8. Octet exceptions in B, Be, H
9. **Bond order ↑ → bond length ↓, energy ↑**
10. Water is **polar** even though O–H is covalent

Past NET MCQs:

Q1:
Which compound has a coordinate covalent bond?
A. H_2O **B.** NH_4^+ **C.** CH_4 **D.** HCl
✅ **Correct: B**

Q2:
What is the bond angle in CH_4?
A. 120° **B.** 90° **C.** 109.5° **D.** 180°
✅ **Correct: C**

Q3:
Which bond is most polar?
A. H–Cl **B.** H–F **C.** H–I **D.** H–Br
✅ **Correct: B**

Q4:
How many sigma and pi bonds in C_2H_4 (ethene)?
A. 5σ, 1π **B.** 4σ, 2π **C.** 3σ, 3π **D.** 6σ
✅ **Correct: A**

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Chapter 4: States of Matter- Gases and Liquids

This chapter covers the **behavior of gases**, **ideal vs real gases**, **gas laws**, **intermolecular forces**, and **liquid properties**. These are often tested in NET with **conceptual + formula-based MCQs**.

Section 1: Properties of Gases

🔷 1. Characteristics of Gases

- Highly compressible
- No fixed shape or volume
- Exert pressure uniformly
- Low density
- Mix evenly and completely

🔷 2. Basic Units & Conversions

Quantity	Units
Pressure (P)	atm, mmHg, torr, Pa
Volume (V)	dm^3, cm^3, L
Temperature (T)	K (Kelvin) = °C + 273
Amount (n)	mole

Conversion:
1 atm = 760 mmHg = 101325 Pa
1 dm^3 = 1000 cm^3 = 1 L

◆ 3. Gas Laws

Law	Formula	Relationship
Boyle's Law	**$P_1V_1 = P_2V_2$**	$P \propto 1/V$ (T constant)
Charles' Law	**$V_1/T_1 = V_2/T_2$**	$V \propto T$ (P constant)
Gay-Lussac's Law	**$P_1/T_1 = P_2/T_2$**	$P \propto T$ (V constant)
Avogadro's Law	**$V \propto n$**	Equal volume → equal moles
Combined Law	**$P_1V_1/T_1 = P_2V_2/T_2$**	Combines all above
Ideal Gas Law	**$PV = nRT$**	R = 0.0821 atm·dm³/mol·K or 8.314 J/mol·K

◆ 4. Dalton's Law of Partial Pressures

$P_{total} = P_1 + P_2 + P_3 + \ldots$

- Useful in calculating partial pressure in gas mixtures.

◆ 5. Graham's Law of Diffusion

$Rate_1 \div Rate_2 = \sqrt{(M_2 \div M_1)}$

- Lighter gases diffuse faster.

🔷 6. Real Gases vs Ideal Gases

Feature	Ideal Gas	Real Gas
Volume of particles	Neglected	Considered
Intermolecular forces	Absent	Present
Behavior	Obeys PV=nRT	Deviates under high P or low T

🔷 7. Van der Waals Equation

$(P + a(n^2/V^2))\ (V - nb) = nRT$

- Corrects for **intermolecular forces (a)** and **volume (b)** in real gases.

Section 2: Liquids and Intermolecular Forces:

🔷 1. Intermolecular Forces

Force Type	Description	Strength
London Dispersion	Instantaneous dipoles	Weakest
Dipole-Dipole	Polar molecules	Moderate
Hydrogen Bonding	H with N, O, or F	Strongest

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2. Properties of Liquids

Property	Definition
Surface Tension	Force at liquid surface (↑ with ↑ IMF)
Viscosity	Resistance to flow (↑ with ↑ IMF)
Vapor Pressure	Pressure of evaporated molecules
Boiling Point	Temp where vapor pressure = external pressure

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3. Boiling Point Trends

- **↑ IMF = ↑ Boiling Point**
- H_2O has high boiling point due to **hydrogen bonding**

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4. Anomalies of Water

- Water expands on freezing
- Ice is less dense than liquid water
- High specific heat capacity

Important Formulas to Remember:

1. **$PV = nRT$**
2. **$P_1V_1/T_1 = P_2V_2/T_2$**
3. **Rate of diffusion $\propto 1/\sqrt{M}$**
4. **$M = dRT/P$** (for molar mass using gas density)
5. **$P_{total} = P_{gas} + P_{H_2O}$ (if over water)**
6. **$K = °C + 273$**
7. **1 mol gas = 22.4 dm^3 at STP**

NET MCQ Tips:

1. If **temperature** increases, **volume** increases (Charles)
2. Always **convert temp to Kelvin** in gas laws
3. **Use PV = nRT** only when 3 quantities are known
4. **Hydrogen bonding** always increases b.p. (common trap)
5. For **rate of diffusion**, lightest gas = fastest
6. Use **Dalton’s Law** in gas mixture problems
7. High pressure & low temp = **real gases deviate** most
8. Remember **1 mole of any gas = 22.4 dm^3 at STP**

Past NET MCQs:

Q1:

If 2.5 moles of gas occupy 5.6 dm^3 at 273 K, what is the pressure? (R = 0.0821)
A. 1 atm **B.** 10 atm **C.** 2 atm **D.** 5 atm

✅ **Correct: B**

→ $P = nRT/V = (2.5)(0.0821)(273) \div 5.6 \approx 10$ atm

Q2:

Which force is strongest among the following?
A. London Forces **B.** Hydrogen Bonding **C.** Dipole-dipole **D.** van der Waals

✅ **Correct: B**

Q3:

Which gas diffuses faster: O_2 or H_2?
A. O_2 **B.** H_2 **C.** Both same **D.** None

✅ **Correct: B**

→ H_2 is lighter than O_2

Q4:
At STP, volume of 3 moles of gas is:
A. 67.2 dm^3 **B.** 22.4 dm^3 **C.** 33.6 dm^3 **D.** 44.8 dm^3

✅ **Correct: A**

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Chapter 5: Chemical Thermodynamics

Core Concepts:

Thermodynamics deals with the **energy changes** (especially heat) during chemical reactions. It includes laws, enthalpy, internal energy, spontaneity, and heat transfer — all **important topics for NET MCQs**.

🔷 1. System and Surroundings

Term	Description
System	Part of universe under study
Surroundings	Everything else
Open System	Exchanges mass & energy (e.g., open cup)
Closed System	Exchanges energy but not mass (e.g., sealed container)
Isolated System	No exchange of mass or energy (e.g., thermos)

🔷 2. Types of Thermodynamic Processes

Type	Constant
Isothermal	Temperature
Adiabatic	No heat exchange ($q = 0$)
Isochoric	Volume
Isobaric	Pressure

◆ 3. First Law of Thermodynamics

$\Delta U = q + w$

- **ΔU:** Change in internal energy
- **q:** Heat added to the system (+ve if absorbed)
- **w:** Work done on the system (+ve if compressed)

For gases:
$w = -P\Delta V$ → expansion = –ve work

◆ 4. Enthalpy (H)

$H = U + PV$

- Change in enthalpy at constant pressure:
  $\Delta H = \Delta U + P\Delta V$

Reaction	ΔH Value
Endothermic	$\Delta H > 0$ (absorbs heat)
Exothermic	$\Delta H < 0$ (releases heat)

Example:
Combustion of methane: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$; $\Delta H = -890$ kJ/mol (exothermic)

◆ 5. Heat Capacity & Calorimetry

- **Specific Heat (c):** Heat needed to raise 1g by 1°C
- **Heat = $q = mc\Delta T$**
  - m = mass, c = specific heat, ΔT = temperature change

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6. Hess's Law

If a reaction occurs in steps, the total enthalpy change is the **sum of individual steps**.

Example:
$C + O_2 \rightarrow CO_2$ $\Delta H = -393$ kJ/mol
$C + \frac{1}{2}O_2 \rightarrow CO$ $\Delta H = -110$ kJ/mol
$CO + \frac{1}{2}O_2 \rightarrow CO_2$ $\Delta H = -283$ kJ/mol

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7. Standard Enthalpies

Quantity	Symbol	Unit
Formation	$\Delta Hf°$	kJ/mol
Combustion	$\Delta Hc°$	kJ/mol
Neutralization	$\Delta Hn°$	kJ/mol

Standard conditions = 298K, 1 atm, 1M

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8. Second Law of Thermodynamics

- **Entropy (S):** Degree of disorder or randomness
  • More disorder = more entropy
  • Gases > liquids > solids
- **$\Delta S > 0$** → spontaneous disorder increases

🔷 **9. Gibbs Free Energy (G)**

$\Delta G = \Delta H - T\Delta S$

ΔG	**Process**
< 0	Spontaneous
> 0	Non-spontaneous
= 0	Equilibrium

Important Equations:

Concept	**Equation**
Internal energy	**$\Delta U = q + w$**
Work done by gas	**$w = -P\Delta V$**
Enthalpy	**$H = U + PV$**
ΔH at constant P	**$\Delta H = \Delta U + P\Delta V$**
Heat absorbed	**$q = mc\Delta T$**
Hess's law	**$\Delta H\_total = \Delta H_1 + \Delta H_2 + \dots$**
Gibbs free energy	**$\Delta G = \Delta H - T\Delta S$**

NET MCQ Tips:

1. **$q = mc\Delta T$** used in calorimetry questions
2. $\Delta H > 0 \rightarrow$ endothermic (e.g., melting, vaporization)
3. $\Delta H < 0 \rightarrow$ exothermic (e.g., combustion)
4. **Hess's Law** is additive: flip and multiply equations carefully
5. Gibbs free energy is used to check **spontaneity**
6. If $\Delta G < 0$, the process occurs **without external input**
7. **Standard enthalpies** are used for $\Delta H°$ calculations
8. Entropy increases in **melting, vaporization, mixing gases**
9. Always convert temp to **Kelvin** in thermodynamic formulas
10. **Work is negative** when gas expands

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Past NET MCQs:

Q1:
If 100g of water is heated from 25°C to 75°C, how much heat is required? (c = 4.18 J/g·K)
A. 20.9 kJ **B.** 10.5 kJ **C.** 15.6 kJ **D.** 5.2 kJ
✅ **Correct: A**
$q = mc\Delta T = 100 \times 4.18 \times (75 - 25) = 20900$ J = **20.9 kJ**

Q2:
Which of the following has highest entropy?
A. Ice **B.** Steam **C.** Water **D.** Diamond
✅ **Correct: B**

Q3:
In which process is ΔH negative?
A. Melting **B.** Boiling **C.** Combustion **D.** Vaporization
✅ **Correct: C**

Q4:
If $\Delta H = -200$ kJ/mol, $\Delta S = -0.4$ kJ/mol·K, what is ΔG at 300K?
A. –80 kJ **B.** –100 kJ **C.** –320 kJ **D.** –200 kJ
✅ **Correct: A**
$\Delta G = \Delta H - T\Delta S = -200 - (300 \times -0.4) = -200 + 120 = -80$ kJ

Chapter 6: Chemical Equilibrium

This chapter explains how chemical reactions reach a dynamic balance and how we can **predict the direction**, **extent**, and **conditions favoring products or reactants** — making it very important for NET.

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1. Reversible Reactions:

- **Reversible Reaction:** Occurs in both forward and reverse directions
  $A + B \rightleftharpoons C + D$
- At equilibrium, the **rate of forward reaction = rate of reverse reaction**
- Concentration remains **constant** (but not equal)

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2. Law of Mass Action:

For a general reaction:
$aA + bB \rightleftharpoons cC + dD$

Equilibrium constant (Kc):

Kc=[C]c[D]d[A]a[B]bK_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}Kc =[A]a[B]b[C]c[D]d

- Only includes **aqueous** and **gaseous** species
- **Solids and pure liquids are excluded**

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3. Characteristics of Equilibrium

Feature	Description
Dynamic	Forward = Reverse reaction rates
Closed System	Required to reach equilibrium
Constant Concentrations	No change after equilibrium
Affected by conditions	T, P, concentration affect position

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4. Types of Equilibrium Constants:

Symbol	Description
Kc	Based on concentration (mol/dm^3)
Kp	Based on partial pressure (atm)
Kp = Kc(RT)^Δn	Δn = moles (products – reactants)

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5. Reaction Quotient (Qc):

Used to **compare current concentrations** with equilibrium:

Qc=[products][reactants]Q_c = \frac{[products]}{[reactants]}Qc =[reactants][products]

Comparison	Direction
Q < K	Forward
Q = K	Equilibrium
Q > K	Reverse

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6. Le Chatelier's Principle:

"If a system at equilibrium is disturbed, it will shift to oppose the change."

Stress	System Shift
Add Reactant	→ (Right)
Remove Reactant	← (Left)
Add Product	←
Remove Product	→
↑ Pressure (gases only)	Toward fewer moles
↑ Temperature	Favors **endothermic** direction
↓ Temperature	Favors **exothermic** direction

Example:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ + heat
• ↑T → shifts left (endothermic)
• ↑P → shifts right (fewer moles)

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7. Factors Affecting Kc:

- **Temperature affects value of Kc**
- **Catalyst does not affect Kc**, only speeds up equilibrium

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8. Important Formulas:

Concept	Formula
Kc expression	Kc=[products][reactants]K_c = \frac{[products]}{[reactants]}Kc =[reactants][products]
Kp relation	Kp=Kc(RT)ΔnK_p = K_c(RT)^{\Delta n}Kp=Kc(RT)Δn
Qc	Qc=[products][reactants]Q_c = \frac{[products]}{[reactants]}Qc =[reactants][products]
Δn (gas moles)	Δn = mol(products) – mol(reactants)

Graphical Overview:

- Time vs Concentration Graph (Forward & Reverse rates)
- Shift left/right due to T, P, Concentration changes
- Equilibrium concentration plateaus

NET MCQ Tips:

1. Kc expression includes **gases and aqueous**, not solids
2. Δn must be calculated **only for gases** in Kp = Kc(RT)^Δn
3. Adding reactants always shifts equilibrium **to right**
4. **Catalyst** does not shift equilibrium
5. Exothermic reactions: **heat is a product**
6. If Kc is large (>1000), **products favored**
7. If Kc is small (<0.001), **reactants favored**
8. For Qc vs Kc, determine if system moves left or right

9. Always convert **T to Kelvin**
10. **Le Chatelier** is key in conceptual MCQs

Past NET MCQs:

Q1:
Which factor changes the value of Kc?
A. Catalyst **B.** Concentration **C.** Temperature **D.** Pressure
✅ **Correct: C**

Q2:
If Kc > 1, which is favored?
A. Reactants **B.** Products **C.** Both equally **D.** None
✅ **Correct: B**

Q3:
For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$, what happens if pressure is increased?
A. No effect **B.** Shifts left **C.** Shifts right **D.** Kc increases
✅ **Correct: C**

Q4:
If Qc < Kc, the reaction will:
A. Move forward **B.** Be at equilibrium **C.** Move backward **D.** Stop
✅ **Correct: A**

Chapter 7: Acids, Bases, and Salts

This chapter focuses on **acid-base theories**, **strengths**, **pH calculations**, **indicators**, and **salt formation** — all commonly tested in NET **theory and calculation-based MCQs**.

---

1. Definitions of Acids & Bases:

Theory	Acid	Base
Arrhenius	Produces H^+ in water	Produces OH^- in water
Brønsted–Lowry	Donates a proton (H^+)	Accepts a proton (H^+)
Lewis	Electron pair acceptor	Electron pair donor

Example:

- HCl is an acid (donates H^+)
- NH_3 is a base (accepts H^+)

---

2. Conjugate Acid-Base Pairs:

Every acid has a **conjugate base** (after donating H^+), and every base has a **conjugate acid** (after accepting H^+).

Example:

- HCl (acid) → Cl^- (conjugate base)
- NH_3 (base) → NH_4^+ (conjugate acid)

Strong vs Weak Acids/Bases:

Strength	Behavior	Example
Strong Acid/Base	Completely ionizes	HCl, HNO_3, NaOH
Weak Acid/Base	Partially ionizes	CH_3COOH, NH_3

---

4. pH and pOH:

Type	pH Range
Acidic	0–6.9
Neutral	7
Basic	7.1–14

Tip: Strong acids/bases give **very low or high pH**.

---

5. Ionization Constants:

- **Ka (acid dissociation constant):**

Ka=[H+][A−][HA]Ka = \frac{[H^+][A^-]}{[HA]}Ka=[HA][H+][A−]

- **Kb (base dissociation constant):**

Kb=[BH+][OH−][B]Kb = \frac{[BH^+][OH^-]}{[B]}Kb=[B][BH+][OH−]

- **Water Ion Product:**

Kw=[H+][OH−]=1×10−14 at 25°CKw = [H^+][OH^-] = 1 × 10^{-14} \text{ at 25°C}Kw=[H+][OH−]=1×10−14 at 25°C

Important Relation:

pKa+pKb=14pKa + pKb = 14pKa+pKb=14

---

6. Salt Formation:

- **Acid + Base → Salt + Water**

Types of salts based on parent acid/base:

- **Neutral Salt:** Strong acid + strong base (e.g., $NaCl$)
- **Acidic Salt:** Strong acid + weak base (e.g., NH_4Cl)
- **Basic Salt:** Weak acid + strong base (e.g., CH_3COONa)

---

7. Acid-Base Indicators:

Indicator	Acid Color	Base Color	pH Range
Litmus	Red	Blue	4.5 – 8.3
Phenolphthalein	Colorless	Pink	8.2 – 10
Methyl Orange	Red	Yellow	3.1 – 4.4

Use indicators to identify equivalent points in titrations.

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Important Equations:

Concept	Formula
pH	**$pH = -\log[H^+]$**
pOH	**$pOH = -\log[OH^-]$**
pH + pOH	**= 14**
Kw	**$Kw = [H^+][OH^-] = 1 \times 10^{-14}$**
Ka/Kb	Ka=[H+][A−][HA]Ka = \frac{[H^+][A$^-$]}{[HA]}Ka=[HA][H+][A−], Kb=[BH+][OH−][B]Kb = \frac{[BH^+][OH$^-$]}{[B]}Kb=[B][BH+][OH−]
pKa + pKb	**= 14**

NET MCQ Tips:

1. Strong acids fully ionize → pH < 2
2. Weak acids have higher pH (around 4–6)
3. **$Kw = 10^{-14}$** always — use for $[H^+]$ and $[OH^-]$
4. Acidic salt = strong acid + weak base
5. pH + pOH = 14 holds only at 25°C
6. If given $[OH^-]$, calculate pOH first
7. Indicators change color over **specific pH range**
8. Strong base (e.g., NaOH) has pH close to **14**
9. Titration involves neutralization; endpoint shown by **indicator color change**
10. Use **log rules**: if $[H^+] = 1 \times 10^{-5}$, pH = 5

Past NET MCQs

Q1:
What is the pH of 1×10^{-3} M HCl?
A. 3 **B.** 11 **C.** 7 **D.** 1
✅ **Correct: A**
$pH = -\log[H^+] = -\log(10^{-3}) = 3$

Q2:
Which salt is basic in nature?
A. NH_4Cl **B.** NaCl **C.** CH_3COONa **D.** HNO_3
✅ **Correct: C**

Q3:
What is the color of phenolphthalein in base?
A. Pink **B.** Yellow **C.** Colorless **D.** Red
✅ **Correct: A**

Q4:
Which relation is correct?
A. $pH = \log[OH^-]$ **B.** $K_w = K_a \times K_b$ **C.** $K_a = 1 \times 10^7$ **D.** $pH + pOH = 7$
✅ **Correct: B**

Chapter 8: Solutions

---

1. Basic Definitions:

- **Solution**: A homogeneous mixture of two or more substances.
- **Solute**: The substance that is dissolved (usually in smaller quantity).
- **Solvent**: The substance that dissolves the solute (usually in larger quantity).

---

2. Types of Solutions:

Solute State	Solvent State	Example
Solid	Liquid	Sugar in water
Gas	Liquid	Carbon dioxide in soda
Liquid	Liquid	Alcohol in water
Gas	Gas	Oxygen in nitrogen (air)

---

3. Units of Concentration:

a. Molarity (M)
M = Moles of solute ÷ Volume of solution in dm^3

b. Molality (m)
m = Moles of solute ÷ Mass of solvent in kg

c. Mole Fraction (X)
X = Moles of one component ÷ Total moles of all components

d. Mass Percentage (% w/w)
% w/w = (Mass of solute ÷ Mass of solution) × 100

e. Parts per million (ppm)
ppm = (Mass of solute ÷ Mass of solution) × 1,000,000

---

4. Factors Affecting Solubility:

- **Temperature**: Solubility of most solids increases with temperature.
- **Pressure**: Solubility of gases increases with pressure.
- **Nature of solute and solvent**: "Like dissolves like" (polar solutes dissolve in polar solvents).

---

5. Henry's Law (Solubility of Gases):

The solubility of a gas is directly proportional to the pressure of the gas above the solution.

Formula:
$C = k \times P$
Where:
C = Concentration of gas
k = Henry's constant
P = Partial pressure of gas

---

6. Raoult's Law (Vapor Pressure of Solutions):

The partial vapor pressure of each component in a solution is equal to the mole fraction of the component multiplied by the vapor pressure of the pure component.

Formula:
$P_a = X_a \times P_a^\circ$
Where:
P_a = Partial vapor pressure of component A
X_a = Mole fraction of component A
P_a° = Vapor pressure of pure A

Total vapor pressure of the solution:
$P_{total} = P_a + P_b = X_a \times P_a^\circ + X_b \times P_b^\circ$

---

7. Colligative Properties:

Colligative properties depend on the **number** of solute particles in solution, not their identity.

---

a. Relative Lowering of Vapor Pressure
$(P^\circ - P) \div P^\circ = X_{solute}$

---

b. Boiling Point Elevation
$\Delta T_b = K_b \times m$
Where:
ΔT_b = Increase in boiling point
K_b = Boiling point elevation constant
m = Molality of solution

---

c. Freezing Point Depression
$\Delta T_f = K_f \times m$
Where:
ΔT_f = Decrease in freezing point
K_f = Freezing point depression constant
m = Molality of solution

---

d. Osmotic Pressure (π)
$\pi = C \times R \times T$
Where:
π = Osmotic pressure (in atm)
C = Molarity of solution

R = Gas constant = 0.0821 $dm^3 \cdot atm \cdot mol^{-1} \cdot K^{-1}$
T = Temperature in Kelvin

---

8. Summary of Important Formulas

Concept	**Formula**
Molarity	M = Moles ÷ Volume in dm^3
Molality	m = Moles ÷ Mass in kg
Mole Fraction	X = Moles of component ÷ Total moles
Mass Percent	% w/w = (Mass solute ÷ Mass solution) × 100
Ppm	ppm = (Mass solute ÷ Mass solution) × 1,000,000
Henry's Law	C = k × P
Raoult's Law	$P_a = X_a \times P_a°$
Boiling Point Elevation	Delta T_b = K_b × m
Freezing Point Depression	Delta T_f = K_f × m
Osmotic Pressure	π = C × R × T

---

9. NET MCQ Tips:

1. Molarity is used when solution **volume** is given.
2. Molality is used when **temperature-dependent** properties are involved.
3. Use **molality** for boiling/freezing point changes (it is independent of temperature).
4. Ionic compounds dissociate and **increase colligative effect**.
5. Henry's Law applies only to **gases in liquids**.

6. Raoult's Law applies to **ideal solutions**.
7. Colligative properties increase with the **number of solute particles**.
8. Osmotic pressure increases with **concentration and temperature**.
9. For dilute aqueous solutions, molarity ≈ molality.
10. ppm is used for **very dilute** solutions (e.g., impurities in water).

---

10. Past NET MCQ Examples:

Q1: What is the molarity of a solution with 2 moles of NaCl in 1 dm^3 of solution?
A. 0.5 M B. 1.0 M C. 2.0 M D. 4.0 M
✅ Correct Answer: C

Q2: What is the unit of osmotic pressure?
A. mol^{-1} B. atm C. m/s^2 D. dm^3
✅ Correct Answer: B

Q3: Which colligative property is used to determine molar mass?
A. Osmotic pressure B. pH C. Surface tension D. Density
✅ Correct Answer: A

Chapter 9: Electrochemistry

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1. Basic Concepts:

- **Electrochemistry** is the study of the relationship between **electricity and chemical reactions**.
- It involves **redox reactions**, where **oxidation** and **reduction** occur.

---

2. Oxidation and Reduction:

- **Oxidation**: Loss of electrons
- **Reduction**: Gain of electrons
- **Oxidizing Agent**: Accepts electrons (gets reduced)
- **Reducing Agent**: Donates electrons (gets oxidized)

---

3. Electrochemical Cells

There are two main types:

Cell Type	Description
Galvanic (Voltaic)	Converts chemical energy to electrical energy (spontaneous reaction)
Electrolytic	Converts electrical energy to chemical energy (non-spontaneous reaction)

---

4. Galvanic (Voltaic) Cell:

Features:

- Two electrodes: **Anode** (oxidation), **Cathode** (reduction)
- Electrons flow from **anode to cathode**
- Salt bridge allows ion flow to maintain neutrality

Cell Notation Example:

Zn (s) | Zn^{2+} (aq) || Cu^{2+} (aq) | Cu (s)

- Left side: **Anode** (oxidation)
- Right side: **Cathode** (reduction)

---

5. Electrolytic Cell:

- Electrical energy is applied to drive a **non-spontaneous** reaction.
- Electrodes are usually **in the same container**.
- **Anode** is still oxidation, **cathode** is reduction, but **electron flow is forced**.

---

6. Standard Electrode Potential (E°):

- E° = Tendency of a species to gain electrons (measured in volts)
- Measured under **standard conditions**:
  1 M solution, 1 atm pressure, 25°C

---

7. Cell EMF (Electromotive Force):

Formula:
E_cell = E°_cathode – E°_anode

- If E_cell is **positive** → reaction is **spontaneous**
- If E_cell is **negative** → reaction is **non-spontaneous**

---

8. Electrolysis:

- Process of **breaking down** a compound using electricity
- Common examples:
    - Electrolysis of water → H_2 and O_2
    - Electrolysis of molten NaCl → Na and Cl_2

---

9. Faraday's Laws of Electrolysis:

First Law:
Mass of substance deposited ∝ quantity of electricity passed.

Formula:
$m = (Z \times Q)$
Where:
m = mass deposited
Z = electrochemical equivalent
Q = total charge ($Q = I \times t$)

Second Law:
When same charge is passed through different electrolytes,
mass deposited ∝ equivalent weight

---

10. Important Formulas:

Concept	Formula
Cell potential (EMF)	E_cell = E°_cathode – E°_anode
Charge (Coulombs)	Q = I × t
Faraday constant	F = 96500 C/mol e^-
Mass deposited	m = (M × I × t) ÷ (n × F)
Number of Faradays	Q ÷ F
Electrochemical equivalent	Z = m ÷ Q

---

11. NET MCQ Tips:

1. Electrons always flow from **anode to cathode** in galvanic cells.
2. **Oxidation is loss**, **reduction is gain** (OIL RIG).
3. Positive E_cell indicates **spontaneous** reaction.
4. **Salt bridge** maintains charge balance by ion movement.
5. E° values can be used to predict **reaction feasibility**.
6. **Anode is negative** in galvanic cells, **positive** in electrolytic.
7. In electrolysis, **cations go to cathode**, **anions to anode**.
8. **Faraday's constant = 96500 C/mol**
9. Electrolytic cells require **external power source**.
10. The more positive E° = stronger oxidizing agent.

---

12. Past NET MCQs:

Q1:
What is the unit of electric charge?
A. Ohm B. Coulomb C. Volt D. Ampere
✅ Correct Answer: B

Q2:
In a galvanic cell, oxidation occurs at:
A. Cathode B. Both electrodes C. Anode D. Salt bridge
✅ Correct Answer: C

Q3:
Which species is reduced in the following reaction?
$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
A. Cu^{2+} B. Zn C. Cu D. Zn^{2+}
✅ Correct Answer: A

Q4:
E° of Zn = -0.76 V, E° of Cu = +0.34 V. What is E_cell?
A. -1.10 V B. +0.42 V C. +1.10 V D. -0.42 V
✅ Correct Answer: C
(Solution: E_cell = 0.34 – (–0.76) = 1.10 V)

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Chapter 10: Periodic Table and Periodicity

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1. Introduction

The **Periodic Table** is a systematic arrangement of elements in order of increasing atomic number. It reveals trends in chemical and physical properties called **periodic properties**.

---

2. Basic Structure

- **Periods**: Horizontal rows (7 total)
- **Groups**: Vertical columns (18 total)
- **Metals**: Left side
- **Nonmetals**: Right side
- **Metalloids**: Stair-step line between metals and nonmetals

---

3. Key Periodic Properties and Trends

---

a. Atomic Radius

- Definition: Distance from nucleus to outermost electron.
- **Increases down a group**, **decreases across a period**

---

b. Ionization Energy (IE)

- Definition: Energy required to remove an electron from a gaseous atom.
- **Decreases down a group**, **increases across a period**

---

c. Electron Affinity (EA)

- Definition: Energy released when an atom gains an electron.
- **Decreases down a group**, **increases across a period**

---

d. Electronegativity (EN)

- Definition: Ability of an atom to attract electrons in a bond.
- **Decreases down a group**, **increases across a period**

---

e. Metallic and Non-Metallic Character

- **Metallic character increases** down a group
- **Non-metallic character increases** across a period

---

4. Important Terms and Definitions

- **Valency**: Number of electrons an atom can lose, gain, or share.
- **Isoelectronic Species**: Different atoms/ions having the same number of electrons (e.g., O^{2-}, F^-, Ne).
- **Shielding Effect**: Inner electrons reduce attraction between nucleus and outer electrons.
- **Effective Nuclear Charge (Z_{eff})**: Net positive charge experienced by outer electrons.

Formula:
$Z_{eff} = Z - S$
Where:
Z = Atomic number
S = Shielding constant

---

5. Summary Table of Periodic Trends

Property	Down a Group	Across a Period
Atomic Radius	Increases	Decreases
Ionization Energy	Decreases	Increases
Electron Affinity	Decreases	Increases
Electronegativity	Decreases	Increases
Metallic Character	Increases	Decreases
Non-metallic Nature	Decreases	Increases

---

6. NET MCQ Tips

1. **Group 1 (Alkali metals)** have the **largest atomic radii** in each period.
2. Noble gases have **zero electron affinity** (they are stable).
3. **Fluorine** is the most electronegative element (EN = 4.0).
4. **Shielding effect** increases with more inner shells.
5. **Ionization energy** is highest for noble gases.
6. Isoelectronic ions can be ranked by atomic number to find size.
7. **Transition elements** show variable oxidation states and incomplete d-subshells.
8. **Effective nuclear charge increases** across a period.
9. **Successive ionization energies** always increase.
10. **Cations are smaller**, **anions are larger** than their parent atoms.

---

7. Important Formulas (Word Format)

- Z_eff = Z – S
- EA = Energy of neutral atom – Energy of anion
- EN (qualitative trend): F > O > N > Cl > S > C > H

---

8. NET Past MCQs

Q1: Which group contains elements with the largest atomic size in a period?
A. Group 1 B. Group 17 C. Group 18 D. Group 13
✅ Correct: A

Q2: Which of the following has the smallest atomic radius?
A. Na B. Mg C. Al D. Cl
✅ Correct: D

Q3: The most electronegative element is:
A. Oxygen B. Fluorine C. Nitrogen D. Chlorine
✅ Correct: B

Q4: Ionization energy increases across a period due to:
A. More shielding
B. Less nuclear charge
C. Smaller atomic radius
D. More shells
✅ Correct: C

Chapter 11: Transition Elements

---

1. Introduction:

- **Transition elements** are the elements found in **Groups 3 to 12** of the periodic table.
- They are also called **d-block elements** because their **d-orbitals** are being filled.
- These elements form a **bridge** between the s-block and p-block elements.

---

2. General Properties of Transition Elements:

1. **Variable oxidation states**
   - e.g., Fe^{2+} and Fe^{3+}, Cu^{+} and Cu^{2+}
2. **Colored compounds**
   - Due to **d-d transitions** (electrons jumping between d-orbitals)
3. **Paramagnetism**
   - Unpaired electrons in d-orbitals cause magnetic behavior
4. **Catalytic activity**
   - They act as **catalysts** (e.g., Fe in Haber process)
5. **Formation of complex compounds**
   - Transition metals can form **coordination complexes**
6. **High melting and boiling points**
7. **Hard and dense metals**

---

3. Electronic Configuration:

General electronic configuration:
$(n - 1)d^{1-10}\ ns^{1-2}$
Example:

- Sc (Z = 21): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$
- Fe (Z = 26): $[Ar]\ 3d^6\ 4s^2$

---

4. Oxidation States of Selected Transition Elements

Element	Common Oxidation States
Sc	+3
Ti	+3, +4
V	+2, +3, +4, +5
Cr	+2, +3, +6
Mn	+2, +4, +7
Fe	+2, +3
Co	+2, +3
Ni	+2
Cu	+1, +2
Zn	+2 (not a true transition metal)

---

5. Colored Compounds

Color arises from **electron transitions between split d-orbitals** under influence of ligands (crystal field theory).

Ion	**Color**
Cu^{2+}	Blue
Cr^{3+}	Green
Fe^{3+}	Yellow/Brown
MnO_4^-	Purple
Ni^{2+}	Green

---

6. Catalytic Properties:

Reaction	**Catalyst**
Haber Process (NH_3 synthesis)	Fe
Contact Process (H_2SO_4 production)	V_2O_5
Hydrogenation of oils	Ni
Decomposition of H_2O_2	MnO_2

---

7. Important Points (NET-Oriented)

- **Transition elements** show **variable oxidation states** due to close energy of 3d and 4s orbitals.
- **d-block elements** form **complex ions** with ligands.
- Most transition metal ions are **colored**.
- **Zn**, though in d-block, is **not** a true transition element (no partially filled d-subshell).
- **Scandium and Zinc** have only **one oxidation state**.
- **Magnetism** depends on the number of **unpaired electrons**.

---

8. NET MCQ Tips:

1. **More unpaired electrons = stronger magnetic property**
2. **d-block = Groups 3 to 12**
3. **Zn, Cd, Hg** are **not true transition metals**
4. Transition elements form **colored ions and complexes**
5. **Fe is used in Haber Process**, **V_2O_5 in Contact Process**
6. Transition metals are good **conductors of heat and electricity**
7. **Catalysts** work by providing an alternate **reaction path with lower activation energy**
8. **Sc^{3+} has no d-electrons** = colorless and nonmagnetic
9. Stability of oxidation states varies across the period
10. Transition metals show **high melting/boiling points** due to strong metallic bonding

---

9. NET Past MCQs

Q1: Which of the following is not a transition element?
A. Fe B. Cu C. Zn D. Cr
✅ Correct Answer: C

Q2: The oxidation state of Mn in $KMnO_4$ is:
A. +2 B. +4 C. +6 D. +7
✅ Correct Answer: D

Q3: Which of the following ions is green in color?
A. Fe^{2+} B. Cr^{3+} C. Mn^{2+} D. Cu^{+}
✅ Correct Answer: B

Q4: Iron is used as a catalyst in:
A. Contact Process B. Haber Process
C. Ostwald Process D. Decomposition of H_2O_2
✅ Correct Answer: B

CHEMISTRY Mastery for NET

Chapter 12: Organic Chemistry – Alkanes, Alkenes, Alkynes

---

1. Introduction:

Organic chemistry deals with **carbon-containing compounds**. The simplest types are **hydrocarbons**, which include:

- **Alkanes**: Single bonds only (saturated)
- **Alkenes**: At least one double bond (unsaturated)
- **Alkynes**: At least one triple bond (unsaturated)

---

2. General Formulas:

Hydrocarbon	General Formula	Example
Alkanes	C_nH_{2n+2}	Methane (CH_4)
Alkenes	C_nH_{2n}	Ethene (C_2H_4)
Alkynes	C_nH_{2n-2}	Ethyne (C_2H_2)

---

3. Nomenclature Basics:

Prefixes for number of carbon atoms:

- 1 = Meth, 2 = Eth, 3 = Prop, 4 = But, 5 = Pent, 6 = Hex, etc.

Suffixes:

- Alkanes: -ane
- Alkenes: -ene
- Alkynes: -yne

Example Names:

- CH_4 = Methane
- CH_3CH_3 = Ethane
- $CH_2{=}CH_2$ = Ethene
- $CH{\equiv}CH$ = Ethyne

---

4. Isomerism:

Structural isomerism is common in alkenes and alkynes.
Alkanes with more than 3 carbon atoms can also form **branched isomers**.

---

5. Reactions of Hydrocarbons:

---

a. Alkanes (Relatively unreactive)

- **Combustion**:
  $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
- **Halogenation** (in presence of UV light):
  $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

---

b. Alkenes (More reactive due to double bond)

- **Hydrogenation**:
  $CH_2{=}CH_2 + H_2 \rightarrow CH_3\text{–}CH_3$ (Nickel catalyst)
- **Halogenation**:
  $CH_2{=}CH_2 + Br_2 \rightarrow CH_2Br\text{–}CH_2Br$
- **Hydration**:
  $CH_2{=}CH_2 + H_2O \rightarrow CH_3\text{–}CH_2OH$ (H^+ catalyst)

---

c. Alkynes

- **Hydrogenation**:
  $HC\equiv CH + 2H_2 \rightarrow CH_3–CH_3$
- **Halogenation**:
  $HC\equiv CH + Br_2 \rightarrow CHBr=CHBr \rightarrow CHBr_2–CHBr_2$

---

6. Physical Properties:

Property	**Alkanes**	**Alkenes**	**Alkynes**
Saturation	Saturated	Unsaturated	Unsaturated
Bond Type	Single (σ)	One π + σ	Two π + σ
Boiling Point	Increases with size	Similar trend	Similar trend
Solubility	Insoluble in water	Insoluble	Insoluble

7.Combustion Comparison

Hydrocarbon	**Complete Combustion**	**Incomplete Combustion**
Alkane	$CO_2 + H_2O$	$CO + H_2O$ or C (soot)
Alkene	$CO_2 + H_2O$	Less efficient; produces CO
Alkyne	$CO_2 + H_2O$	Smoky flame (more soot)

---

8. NET Tips:

1. **Alkanes** undergo substitution, not addition.
2. **Alkenes and alkynes** undergo **addition reactions**.
3. Unsaturation can be tested using **bromine water (turns colorless)**.

4. **Combustion** gives CO_2 and H_2O — highly exothermic.
5. Reactivity: **Alkynes > Alkenes > Alkanes**.
6. **Hydrogenation** converts unsaturated → saturated.
7. Addition of water across double bond = **hydration**.
8. More isomers are possible as **carbon atoms increase**.
9. **Alkanes are nonpolar**, hence insoluble in water.
10. **Bromine water test** is a common NET concept.

---

9. NET Past MCQs:

Q1: Which compound decolorizes bromine water?
A. Ethane B. Ethene C. Methane D. Propane
✅ Correct Answer: B

Q2: What is the general formula of alkenes?
A. C_nH_{2n+2} B. C_nH_{2n} C. C_nH_{2n-2} D. C_nH_{2n-4}
✅ Correct Answer: B

Q3: Which hydrocarbon burns with a smoky flame?
A. Ethene B. Ethyne C. Propane D. Methane
✅ Correct Answer: B

Q4: Hydrogenation of ethene gives:
A. Ethyne B. Ethanol C. Ethane D. Methanal
✅ Correct Answer: C

CHEMISTRY Mastery for NET

Chapter 13: Functional Groups

(Alcohols, Carboxylic Acids, Aldehydes, Ketones, etc.)

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1. Introduction:

Functional groups are specific groups of atoms in molecules that determine the characteristic **chemical reactions** of those molecules.

Each class of organic compound has a specific functional group responsible for its behavior.

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2. Common Functional Groups and Their Structures:

Functional Group	Formula/Structure	Suffix	Example Compound
Alcohol	$–OH$	-ol	Ethanol (CH_3CH_2OH)
Aldehyde	$–CHO$	-al	Ethanal (CH_3CHO)
Ketone	$–CO–$ (within chain)	-one	Propanone (CH_3COCH_3)
Carboxylic Acid	$–COOH$	-oic acid	Ethanoic acid (CH_3COOH)
Ester	$–COOR$	-oate	Methyl ethanoate (CH_3COOCH_3)
Ether	$–O–$	-oxy	Diethyl ether ($CH_3CH_2OCH_2CH_3$)
Amine	$–NH_2$	-amine	Methylamine (CH_3NH_2)
Amide	$–CONH_2$	-amide	Ethanamide (CH_3CONH_2)

3. Properties and Reactions:

a. Alcohols

- **Physical**: Polar, hydrogen bonding, higher boiling point than alkanes
- **Chemical**:
    - Combustion: $CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
    - Oxidation: Alcohol → Aldehyde → Acid

b. Aldehydes

- React with **Tollen's reagent** → silver mirror
- Easily oxidized to carboxylic acids

c. Ketones

- Do not react with Tollen's or Fehling's reagent
- Used as solvents (e.g., acetone)

d. Carboxylic Acids

- Acidic nature due to **–COOH** group
- React with bases to form salts
- Esterification with alcohols:
  $CH_3COOH + CH_3OH \rightarrow CH_3COOCH_3 + H_2O$ (H^+ catalyst)

e. Esters

- Pleasant-smelling (used in perfumes, flavors)
- Formed via esterification (acid + alcohol)

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f. Amines

- Basic due to **lone pair on nitrogen**
- React with acids to form ammonium salts

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4. Summary Table (Functional Groups):

Class	General Formula	Example	Test or Use
Alcohols	R–OH	Ethanol	Sodium metal test, combustion
Aldehydes	R–CHO	Ethanal	Tollen’s reagent (Ag mirror)
Ketones	R–CO–R′	Acetone	No silver mirror
Acids	R–COOH	Acetic acid	Litmus test, esterification
Esters	R–COOR′	Fruity esters	Perfumes, flavors
Amines	$R–NH_2$	Methylamine	Basic in nature

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5. NET Tips:

1. **Tollen’s reagent** distinguishes **aldehydes from ketones**
2. **Alcohol + acid → ester + water** (with H^+ catalyst)
3. **Esters smell sweet** (important NET MCQ)
4. **Amine groups** are basic (due to lone pair)
5. **Oxidation** of alcohol gives aldehyde → then carboxylic acid
6. Ketones **cannot be oxidized easily**, unlike aldehydes

7. Functional group suffixes help in **IUPAC naming**
8. **Ethanol burns with a blue flame**
9. Carboxylic acids turn **blue litmus red**
10. **Esters are used** in making soaps, perfumes, and plasticizers

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6. NET Past MCQs:

Q1: Which functional group is present in ethanol?
A. –COOH B. –CHO C. –OH D. –CO–
✅ Correct Answer: C

Q2: Which test is used to distinguish between aldehydes and ketones?
A. Bromine water B. Tollen's reagent C. $FeCl_3$ D. NaOH
✅ Correct Answer: B

Q3: Reaction of carboxylic acid and alcohol gives:
A. Ether B. Amine C. Ester D. Salt
✅ Correct Answer: C

Q4: Which compound has a fruity smell?
A. Carboxylic acid B. Ketone C. Alcohol D. Ester
✅ Correct Answer: D

CHEMISTRY Mastery for NET

Chapter 14: Macromolecules

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1. What Are Macromolecules?

- **Macromolecules** are **large molecules** composed of thousands of atoms.
- They are **formed by polymerization**—the linking of small units called **monomers**.

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2. Types of Macromolecules:

Macromolecules are mainly categorized into:

Type	**Description**	**Example**
Natural Polymers	Found in nature	Proteins, DNA, starch
Synthetic Polymers	Man-made through chemical processes	Nylon, Teflon, PVC
Semi-synthetic	Modified natural polymers	Rayon (modified cellulose)

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3. Polymerization:

- **Polymerization** is the chemical process by which monomers are joined to form polymers.

a. Addition Polymerization

- Monomers add to each other without elimination of any small molecule.
- **Example**:
  $n(CH_2{=}CH_2) \rightarrow –CH_2–CH_2–_n$ (Polyethene)

b. Condensation Polymerization

- Monomers join with the **elimination** of a small molecule (e.g., H_2O).
- **Example**: Formation of nylon from diamine and dicarboxylic acid.

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4. Important Synthetic Polymers

Polymer	Monomer	Use
Polyethene	Ethene ($CH_2=CH_2$)	Plastic bags, containers
Polypropene	Propene ($CH_2=CHCH_3$)	Furniture, ropes
Polystyrene	Styrene ($C_6H_5CH=CH_2$)	Foam, packaging
PVC	Vinyl chloride ($CH_2=CHCl$)	Pipes, cables
Teflon	Tetrafluoroethene	Non-stick cookware
Nylon-6,6	Hexamethylene diamine + Adipic acid	Fibers, fabrics
Polyester	Dicarboxylic acid + Diol	Textiles, bottles

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5. Natural Macromolecules:

Macromolecule	Monomer	Function
Protein	Amino acids	Growth, repair, enzymes
Carbohydrates	Glucose	Energy storage (starch, glycogen)
Nucleic Acids	Nucleotides (A, T, G, C)	Genetic information
Lipids	Fatty acids + glycerol	Energy storage, insulation

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6. Properties of Polymers:

- Lightweight but strong
- Heat and chemical resistant (e.g., Teflon)
- Non-biodegradable (most synthetics)
- Can be molded easily
- Some are **elastic** (e.g., rubber), some are **rigid**

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7. NET Important Tips:

1. **Addition polymer** = Monomers with **double bonds**
2. **Condensation polymer** = Monomers with **2 functional groups**
3. PVC is made from **vinyl chloride (CH_2=CHCl)**
4. Teflon is chemically resistant and **non-stick**
5. **Natural rubber** is a polymer of **isoprene**
6. Proteins = polymers of **amino acids**
7. Nylon and polyester are **condensation polymers**
8. Polymers are mostly **non-biodegradable**
9. **Biopolymers** like DNA and proteins are **naturally occurring**
10. Polystyrene used in packaging is a **thermoplastic polymer**

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8. NET Past MCQs:

Q1: Which polymer is used in making non-stick cookware?
A. PVC   B. Teflon   C. Nylon   D. Polystyrene
✅ Correct Answer: B

Q2: What type of polymer is nylon?
A. Addition   B. Condensation   C. Co-polymer   D. Thermosetting
✅ Correct Answer: B

Q3: Which of the following is a natural polymer?
A. PVC   B. Protein   C. Polystyrene   D. Nylon
✅ Correct Answer: B

Q4: Which monomers form polyester?
A. Glucose   B. Ethene   C. Dicarboxylic acid and diol   D. Amino acids
✅ Correct Answer: C

CHEMISTRY Mastery for NET

Chapter 15: Environmental Chemistry

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1. Introduction:

Environmental Chemistry deals with the study of chemical processes occurring in the environment, especially those influenced by human activity. It focuses on pollution, its sources, effects, and control methods.

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2. Major Types of Pollution:

Type	Description
Air Pollution	Contamination of air due to gases, particulates, chemicals
Water Pollution	Contamination of water bodies (rivers, lakes, oceans)
Soil Pollution	Degradation of land due to chemicals, waste, pesticides
Noise Pollution	Excessive sound levels disrupting environment and health

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3. Air Pollutants:

a. Primary Pollutants – Directly emitted from sources

- CO (Carbon monoxide)
- NO, NO_2 (Nitrogen oxides)
- SO_2 (Sulfur dioxide)
- Particulate matter (dust, soot)

b. Secondary Pollutants – Formed by reactions in atmosphere

- Ozone (O_3)
- Peroxyacetyl nitrate (PAN)
- Acid rain components (H_2SO_4, HNO_3)

4. Effects of Major Pollutants:

Pollutant	Source	Harmful Effects
CO	Incomplete fuel combustion	Binds with hemoglobin, reduces oxygen in blood
NOx, SOx	Vehicles, industry	Cause acid rain, respiratory issues
Ozone (O_3)	Smog reactions	Eye irritation, lung damage
Lead	Old fuels, paints	Affects nervous system, especially in children
Particulates	Factories, smoke	Asthma, lung diseases

5. Acid Rain:

- **Caused by:** SO_2 and NOx reacting with water vapor
- **Formation:**
  - $SO_2 + O_2 \rightarrow SO_3$
  - $SO_3 + H_2O \rightarrow H_2SO_4$
  - $NO_2 + H_2O \rightarrow HNO_3$
- **Effects:**
  - Destroys crops, forests, aquatic life
  - Corrodes buildings and monuments

6. Greenhouse Effect & Global Warming:

Greenhouse gases (GHGs):
CO_2, CH_4, N_2O, H_2O vapor, O_3, CFCs

- **Greenhouse Effect:** Traps heat in Earth's atmosphere
- **Global Warming:** Rise in Earth's average temperature due to excessive GHGs

Effects:

- Melting of glaciers
- Rising sea levels
- Extreme weather patterns
- Threat to biodiversity

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7. Ozone Layer Depletion:

- **Ozone (O_3)** in stratosphere protects from **UV radiation**
- **CFCs (Chlorofluorocarbons)** destroy ozone

Reaction:
$CFCl_3 \rightarrow Cl\bullet$ + others
$Cl\bullet + O_3 \rightarrow ClO\bullet + O_2$
$ClO\bullet + O_3 \rightarrow Cl\bullet + 2O_2$

- **Effect:** UV rays increase risk of skin cancer, cataracts

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8. Water Pollution

Causes:

- Industrial waste
- Domestic sewage
- Oil spills
- Agricultural runoff (pesticides, fertilizers)

Common Contaminants:

- Nitrates, phosphates
- Pathogens
- Heavy metals (lead, mercury)

Effects:

- Death of aquatic life
- Eutrophication (algal blooms)
- Contaminated drinking water

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9. Environmental Protection:

Preventive Measures:

- **Catalytic converters** in cars to reduce CO and NOx
- **Wastewater treatment plants**
- Use of **bio-degradable products**
- Reduce, reuse, recycle
- International agreements:
  - **Kyoto Protocol** – reduce greenhouse gases
  - **Montreal Protocol** – reduce CFCs

10. NET Tips:

1. CO is **most dangerous** air pollutant due to hemoglobin binding
2. SO_2 and NO_2 cause **acid rain**
3. **Greenhouse gases trap heat**, leading to global warming
4. **CFCs** deplete the ozone layer in the **stratosphere**
5. **BOD** (Biological Oxygen Demand) is a key indicator of water pollution
6. **Eutrophication** is caused by **excess nutrients** in water
7. **Particulate matter** causes **asthma and lung problems**
8. UV rays increase with **ozone depletion**
9. Oil spills are a major cause of **marine pollution**
10. Lead affects the **nervous system**

11. NET Past MCQs:

Q1: Which gas is responsible for global warming?
A. O_2 B. CO C. CO_2 D. Cl_2
✅ Correct Answer: C

Q2: Acid rain is caused by:
A. CO and NO
B. SO_2 and NO_2
C. CO_2 and H_2S
D. CO and O_3
✅ Correct Answer: B

Q3: Which of the following is a greenhouse gas?
A. CH_4 B. SO_2 C. O_2 D. Cl_2
✅ Correct Answer: A

Q4: CFCs are responsible for:
A. Global warming B. Acid rain C. Ozone depletion D. Water pollution
✅ Correct Answer: C