Unit 1: Atomic Structure, Chemical Bonding & Solutions

Engineering Chemistry — DI01000071

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Unit 1 — Overview

Syllabus Topics

1.1 Atomic Structure (Orbits, Orbitals, e^- Config.)
1.2 Chemical Bonding (Ionic, Covalent, Metallic...)
1.3 Types of Solids
1.4 Solutions & Concentration Units
1.5 pH & Buffer Solutions
1.6 Degree of Ionization

Weightage: $\approx 14\%$

Teaching Hrs: ≈ 7 hrs

Section 1.1
Atomic Structure

Orbits vs Orbitals · Shapes of Orbitals · Quantum Numbers
Electronic Configuration · Pauli's Principle
Hund's Rule · Aufbau Principle

Orbits vs Orbitals

Orbit

- Circular/elliptical path
- 2D concept (Bohr model)
- Denoted: $n = 1, 2, 3 \ldots$
- Max. $2n^2$ electrons
- Fixed, well-defined path

Orbital

- 3D probability region
- Quantum mechanical concept
- Denoted: s, p, d, f
- Max. 2 electrons per orbital
- No definite boundary

Key Point

Orbit = outdated fixed path (Bohr); Orbital = modern 3D probability cloud (quantum model)

Shapes of Orbitals

s-Orbital

Spherical

1 per subshell

Max. 2 e^-

p-Orbital

Dumbbell

3 per subshell

Max. 6 e^-

d-Orbital

Double dumbbell

5 per subshell

Max. 10 e^-

Subshell Capacity

Subshell	s	p	d	f
# Orbitals	1	3	5	7
Max e^-	2	6	10	14

Pauli's Exclusion Principle

Statement

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

Four Quantum Numbers

- n — Principal (shell: 1, 2, 3...)
- l — Azimuthal (subshell: 0 to $n-1$)
- m_l — Magnetic (orientation: $-l$ to $+l$)
- m_s — Spin ($+\frac{1}{2}$ or $-\frac{1}{2}$)

Consequence

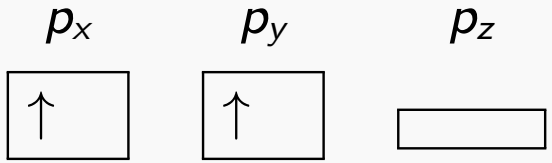
Two electrons in the same orbital must have **opposite spins** ($\uparrow\downarrow$) — max. **2 electrons per orbital**

Hund's Rule of Maximum Multiplicity

Statement

"Electrons occupy each orbital of a subshell *singly* before any pairing begins, with **parallel spins**."

Example: Carbon ($Z = 6$), config $2p^2$



Both electrons have **parallel spins** ($\uparrow\uparrow$) — NOT paired

Key Point

Half-filled (p^3, d^5) and fully-filled (p^6, d^{10}) subshells are **extra stable**

Aufbau Principle

Statement

"Electrons fill orbitals in order of **increasing energy**, governed by the $(n + l)$ rule."

Filling Order

$$1s \to 2s \to 2p \to 3s \to 3p \to 4s \to 3d \to 4p \to 5s \to 4d \to 5p \to 6s \to 4f \to 5d \ldots$$

$(n + l)$ Rule

- Orbital with **lower** $(n + l)$ fills first
- If $(n + l)$ equal, **lower** n fills first
- Ex: $4s$ $(n + l = 4)$ fills before $3d$ $(n + l = 5)$

Electronic Configuration: $Z = 1$ to 10

Z	Element	Configuration	Val. e^-
1	Hydrogen (H)	$1s^1$	1
2	Helium (He)	$1s^2$	2
3	Lithium (Li)	$[He]\,2s^1$	1
4	Beryllium (Be)	$[He]\,2s^2$	2
5	Boron (B)	$[He]\,2s^2\,2p^1$	3
6	Carbon (C)	$[He]\,2s^2\,2p^2$	4
7	Nitrogen (N)	$[He]\,2s^2\,2p^3$	5
8	Oxygen (O)	$[He]\,2s^2\,2p^4$	6
9	Fluorine (F)	$[He]\,2s^2\,2p^5$	7
10	Neon (Ne)	$[He]\,2s^2\,2p^6$	8

Note

He, Ne have **complete valence shells** — noble gas (inert) configuration

Electronic Configuration: $Z = 11$ to 20

Z	Element	Configuration	Val. e^-
11	Sodium (Na)	$[Ne]\,3s^1$	1
12	Magnesium (Mg)	$[Ne]\,3s^2$	2
13	Aluminium (Al)	$[Ne]\,3s^2\,3p^1$	3
14	Silicon (Si)	$[Ne]\,3s^2\,3p^2$	4
15	Phosphorus (P)	$[Ne]\,3s^2\,3p^3$	5
16	Sulphur (S)	$[Ne]\,3s^2\,3p^4$	6
17	Chlorine (Cl)	$[Ne]\,3s^2\,3p^5$	7
18	Argon (Ar)	$[Ne]\,3s^2\,3p^6$	8
19	Potassium (K)	$[Ar]\,4s^1$	1
20	Calcium (Ca)	$[Ar]\,4s^2$	2

Note

After Ar, $4s$ fills **before** $3d$ — Aufbau/$(n + l)$ rule

Electronic Configuration: $Z = 21$–30 (Transition Metals)

Z	Element	Configuration
21	Scandium (Sc)	$[Ar]\,3d^1\,4s^2$
22	Titanium (Ti)	$[Ar]\,3d^2\,4s^2$
23	Vanadium (V)	$[Ar]\,3d^3\,4s^2$
24	Chromium (Cr)	$[Ar]\,3d^{\mathbf{5}}\,4s^{\mathbf{1}}$ (exception)
25	Manganese (Mn)	$[Ar]\,3d^5\,4s^2$
26	Iron (Fe)	$[Ar]\,3d^6\,4s^2$
27	Cobalt (Co)	$[Ar]\,3d^7\,4s^2$
28	Nickel (Ni)	$[Ar]\,3d^8\,4s^2$
29	Copper (Cu)	$[Ar]\,3d^{\mathbf{10}}\,4s^{\mathbf{1}}$ (exception)
30	Zinc (Zn)	$[Ar]\,3d^{10}\,4s^2$

Cr & Cu Exceptions (PYQ – W'23)

Half-filled ($3d^5$) and fully-filled ($3d^{10}$) are **extra stable** — one e^- shifts from $4s \rightarrow 3d$

Section 1.2
Chemical Bonding

Octet Rule · Ionic · Covalent · Coordinate
Metallic · Hydrogen Bond · van der Waals

Octet Rule & Types of Chemical Bonds

Octet Rule

Atoms gain, lose, or share electrons to achieve **8 electrons** in the valence shell (noble gas configuration).

Bond Types

Bond	Mechanism	Example
Ionic	Electron transfer	NaCl, MgO
Covalent	Electron sharing	H_2, CH_4
Coordinate	One-sided e^- donation	NH_4^+, CO
Metallic	Electron sea model	Fe, Cu, Al
Hydrogen	Dipole–dipole (X-H$\cdots$Y)	H_2O, HF, DNA
van der Waals	Induced dipole	Noble gases, CH_4

Ionic Bond

Definition

Formed by complete **transfer of electrons** from metal to non-metal. Resulting oppositely charged ions attract electrostatically.

Example: NaCl (PYQ – Su'23)

$$\text{Na}\ (2,8,\mathbf{1}) \xrightarrow{-e^-} \text{Na}^+\ (2,8)$$

$$\text{Cl}\ (2,8,\mathbf{7}) \xrightarrow{+e^-} \text{Cl}^-\ (2,8,8)$$

$$\Rightarrow \text{Na}^+\text{Cl}^- \text{ — ionic lattice}$$

Properties

- High m.p. & b.p. (strong electrostatic forces)
- Conducts electricity in molten/dissolved state
- Soluble in polar solvents (water); brittle solid

Covalent Bond

Definition

Formed by **mutual sharing of electron pairs** between two non-metal atoms. Each atom contributes electrons.

Types

- **Single**: H–H (1 shared pair)
- **Double**: O=O (2 shared pairs)
- **Triple**: N≡N (3 shared pairs)

Properties

- Low m.p. & b.p.
- Poor conductors
- Non-polar solvent solubility
- Directional bond

PYQ (W'24)

Draw Lewis structure of ethene (C_2H_4); explain σ and π bonds.

Coordinate / Dative Bond

Definition

A covalent bond in which **both electrons are donated by one atom** (donor) to another (acceptor). Shown by arrow: donor $\rightarrow$ acceptor.

NH_4^+ Formation

NH_3 lone pair $\rightarrow H^+$

$\Rightarrow NH_4^+$ (ammonium ion)

All 4 N–H bonds identical

H_3O^+ Formation

H_2O lone pair $\rightarrow H^+$

$\Rightarrow H_3O^+$ (hydronium)

Key in acid-base chemistry

Key Distinction

Coordinate bond = special case of covalent bond; once formed, it is **indistinguishable** from ordinary covalent bond

Metallic Bond

Electron Sea Model

Metal atoms release valence electrons $\rightarrow$ form a **"sea" of delocalized electrons** surrounding a lattice of positive metal ions (kernels).

Properties Explained

Property	Reason
High conductivity	Free electrons move under electric field
Thermal conductivity	Free electrons transfer kinetic energy
Malleability/ductility	Layers slide; bond non-directional
Metallic lustre	Free electrons absorb & re-emit light
High tensile strength	Strong electrostatic attraction of ions to e^- sea

Hydrogen Bond

Definition

Electrostatic attraction between a **hydrogen atom bonded to a highly electronegative atom** (F, O, N) and a lone pair of another electronegative atom. Shown as X–H$\cdots$Y.

Types

Intermolecular:
H_2O, HF, NH_3

Intramolecular:
Salicylaldehyde, o-nitrophenol

Significance

- High b.p. of H_2O (100°C)
- DNA double helix
- Protein folding
- Ice less dense than water

van der Waals Forces

Overview

Weak intermolecular forces arising from **temporary or permanent dipoles**. Weakest of all intermolecular forces.

Three Types

Type	Interaction	Example
London / Dispersion	Instantaneous induced dipole–dipole	Noble gases, CH_4, I_2
Dipole–Dipole	Permanent dipole	HCl, SO_2
Dipole–Induced	Permanent + induced dipole	HCl–Ar mixture

Key Point

Strength order: London $<$ Dipole-Dipole $<$ H-bond $<$ Ionic/Covalent

Bond Comparison (PYQ – W'23, Su'24)

Bond	Mechanism	Strength	Example
Ionic	e^- transfer	Very strong	NaCl
Covalent	Shared pair	Strong	H_2, CH_4
Coordinate	One-sided donation	Strong	NH_4^+, BF_3
Metallic	e^- sea	Strong	Fe, Cu
Hydrogen	Dipole attraction	Moderate	H_2O, HF
van der Waals	Induced dipole	Weak	Noble gases

Exam Tip

A 7-mark question may ask to **compare any 3 bond types** with definition, formation, and properties. Prepare a table answer.

Section 1.3
Types of Solids

Ionic · Covalent / Network · Metallic · Molecular

Properties · Crystal Structures · Engineering Applications

Four Types of Solids

Type	Particles	Bond	Example
Ionic	Cations + anions	Ionic	NaCl, CaF_2
Covalent	Atoms (network)	Covalent	Diamond, SiO_2
Metallic	Metal ions + e^- sea	Metallic	Fe, Cu, Al
Molecular	Molecules	van der Waals, H-bond	Ice, CO_2, I_2

Highest m.p.: Covalent (diamond, 3550 °C)

Lowest m.p.: Molecular (CO_2 sublimes at −78 °C)

Crystal Structures: FCC, BCC, HCP

FCC

Face-Centred Cubic

APF: 74%

CN: 12

Ex: Cu, Al, Ni

BCC

Body-Centred Cubic

APF: 68%

CN: 8

Ex: Fe(α), Cr, W

HCP

Hexagonal Close-Packed

APF: 74%

CN: 12

Ex: Mg, Ti, Zn

APF = Atomic Packing Factor

$$\text{APF} = \frac{\text{Volume of atoms in unit cell}}{\text{Total volume of unit cell}}$$

Simple cubic: 52%

Section 1.4
Solutions & Concentration Units

Molarity · Normality · Molality · Mole Fraction

% by Mass · % by Volume · pH & pOH

Solution Terminology

Term	Meaning
Solute	Substance dissolved (smaller amount)
Solvent	Dissolving medium (larger amount)
Solution	Homogeneous mixture of solute + solvent
Concentration	Amount of solute per unit volume/mass of solution
Dilute solution	Small amount of solute relative to solvent
Concentrated solution	Large amount of solute relative to solvent
Saturated	No more solute can dissolve at given temp.
Unsaturated	Can dissolve more solute
Supersaturated	More dissolved than saturation (unstable)

Molarity (M)

Definition & Formula

Molarity = no. of moles of solute per litre of solution

$$M = \frac{\text{moles of solute}}{\text{volume of solution (L)}} = \frac{w \times 1000}{M_w \times V_{\text{mL}}}$$

Solved Example (PYQ – Su'23)

Find molarity of 4 g NaOH in 500 mL solution.

$M_w(\text{NaOH}) = 40\,\text{g/mol}$

$$M = \frac{4 \times 1000}{40 \times 500} = \mathbf{0.2\,M}$$

Normality (N)

Definition & Formula

Normality = no. of gram equivalents of solute per litre of solution

$$N = \frac{\text{gram equivalents}}{\text{volume (L)}} = \frac{w \times 1000}{E_w \times V_{\text{mL}}}; \quad E_w = \frac{M_w}{n}$$

n-factor (equivalence factor)

Acid/Base	Compound	n-factor
Acid	HCl	1
Acid	H_2SO_4	2
Base	NaOH	1
Base	$Ca(OH)_2$	2

$$N = M \times n$$

Molality & Mole Fraction

Molality (m)

Moles of solute per **kg of solvent** (temperature-independent)

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{w \times 1000}{M_w \times W_s}$$

Mole Fraction (χ)

Ratio of moles of one component to total moles of all components

$$\chi_A = \frac{n_A}{n_A + n_B}; \quad \chi_B = \frac{n_B}{n_A + n_B}; \quad \chi_A + \chi_B = 1$$

Key

Molality & mole fraction are **independent of temperature**; molarity changes with temp.

Percentage Concentration Units

Unit	Formula	Abbr.
% by mass (w/w)	$\frac{w_{\text{solute}}}{w_{\text{soln}}} \times 100$	%(w/w)
% by volume (v/v)	$\frac{V_{\text{solute}}}{V_{\text{soln}}} \times 100$	%(v/v)
% by mass-volume (w/v)	$\frac{w_{\text{solute}}(\text{g})}{V_{\text{soln}}(\text{mL})} \times 100$	%(w/v)

Example

10 g NaCl dissolved in 90 g water:

$$\%(w/w) = \frac{10}{10+90} \times 100 = \mathbf{10\%}$$

pH and pOH

Definitions

$$\text{pH} = -\log[\text{H}^+] \qquad \text{pOH} = -\log[\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14 \quad (\text{at } 25^\circ\text{C})$$

$$K_w = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$$

pH Scale

Acidic (pH < 7)	**Neutral** (pH $= 7$)	**Basic** (pH > 7)
HCl, H_2SO_4	Pure water	NaOH, NH_3

Solved (PYQ – W'24)

$[\text{H}^+] = 0.001\,\text{M}$: $\text{pH} = -\log(10^{-3}) = \mathbf{3}$ (acidic)

pH Calculations — Strong Acids

Strong Acids (fully ionized)

$HA \rightarrow H^+ + A^-$ (100% ionization)

$[H^+] = C \times n$ (n = basicity)

Worked Examples

Solution	$[H^+]$	pH
0.1 M HCl	0.1 M	1.0
0.01 M HCl	0.01 M	2.0
0.1 M H_2SO_4	0.2 M	0.7
1 M NaOH ($\rightarrow$ pOH=0)	10^{-14} M	14.0

Rule

For diprotic acids (H_2SO_4, H_2CO_3): $[H^+] = 2C$ (if fully ionized)

Section 1.5
Degree of Ionization

Arrhenius Theory · Degree of Ionization

Factors Affecting Ionization · Ostwald's Dilution Law

Arrhenius Theory of Ionization

Theory (1887)

Acids, bases, and salts **dissociate into ions** when dissolved in water.

Acid

Produces $\mathbf{H^+}$ ions in water

$HCl \rightarrow H^+ + Cl^-$

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

Base

Produces $\mathbf{OH^-}$ ions in water

$NaOH \rightarrow Na^+ + OH^-$

$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

Limitation

Valid only in **aqueous** solutions. Does not explain acidic/basic behaviour in non-aqueous solvents.

Degree of Ionization (α)

Definition & Formula

Fraction of dissolved electrolyte that ionizes

$$\alpha = \frac{\text{no. of moles ionized}}{\text{total moles dissolved}} \quad (0 \leq \alpha \leq 1)$$

Strong Electrolytes

$\alpha \approx 1$ (100%)

HCl, H_2SO_4,
NaOH, NaCl

Weak Electrolytes

$\alpha \ll 1$ ($< 5\%$)

CH_3COOH,
NH_4OH, HCN

PYQ (Su'24)

Define degree of ionization. State 4 factors affecting it.

Factors Affecting Degree of Ionization

Factor	Effect
Nature of solute	Strong electrolytes ionize $>$ weak electrolytes
Nature of solvent	Polar solvents (water) increase ionization
Concentration	α increases on dilution
Temperature	Higher temp $\Rightarrow$ higher α
Common ion effect	Adding common ion $\Rightarrow$ suppresses ionization
Dielectric constant	Higher ε of solvent $\Rightarrow$ higher α

Ostwald's Dilution Law

$$K_a = \frac{C\alpha^2}{1-\alpha} \approx C\alpha^2 \text{ (when } \alpha \ll 1) \Rightarrow \alpha = \sqrt{K_a/C}$$

Section 1.6
Buffer Solutions

Definition · Types · Buffer Action
Henderson-Hasselbalch Equation · Applications

Buffer Solutions

Definition (PYQ – W'25, Su'25)

A solution that **resists change in pH** upon addition of small amounts of acid or base (or on dilution).

Acidic Buffer

pH $<$ 7

Weak acid + its salt

$CH_3COOH + CH_3COONa$

pH $\approx$ 4.74

Basic Buffer

pH $>$ 7

Weak base + its salt

$NH_4OH + NH_4Cl$

pH $\approx$ 9.25

Buffer Action & Henderson-Hasselbalch Equation

Mechanism (CH_3COOH/CH_3COONa)

On adding acid (H^+):

$CH_3COO^- + H^+ \rightarrow CH_3COOH$ (H^+ consumed)

On adding base (OH^-):

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

Henderson-Hasselbalch Equation

$$pH = pK_a + \log \frac{[Salt]}{[Acid]}$$

$$pOH = pK_b + \log \frac{[Salt]}{[Base]}$$

Applications of Buffer Solutions

Field	Buffer	pH
Blood (biological)	H_2CO_3/HCO_3^-	7.35–7.45
Pharmaceutical/drug	Various	Product-specific
Food preservation	Citric acid/citrate	3–5
Electroplating baths	Acetate/borate buffers	4–9
Analytical chemistry	Standard buffers	4.0, 7.0, 9.2

Blood Buffer (PYQ – W'25)

Blood pH must stay 7.35–7.45. Below 7.35 = **acidosis**; above 7.45 = **alkalosis** — both life-threatening.

Numericals: Molarity & Normality (PYQ – Su'25)

Problem 1

Find molarity of 9.8 g H_2SO_4 in 500 mL solution.

$M_w(H_2SO_4) = 98$, $M = \dfrac{9.8 \times 1000}{98 \times 500} = \mathbf{0.2\,M}$

Problem 2

Find normality of 4.9 g H_2SO_4 in 250 mL solution.

$E_w = 98/2 = 49$; $N = \dfrac{4.9 \times 1000}{49 \times 250} = \mathbf{0.4\,N}$

Relation

$N = M \times n\text{-factor} \quad \Rightarrow 0.4 = 0.2 \times 2$ ✓

Numerical: Buffer pH (PYQ – W'24)

Problem

Calculate pH of acetate buffer: $[CH_3COOH] = 0.1\,M$, $[CH_3COONa] = 0.1\,M$. $K_a(CH_3COOH) = 1.8 \times 10^{-5}$

Solution

$$pK_a = -\log(1.8 \times 10^{-5}) = 4.74$$

$$pH = 4.74 + \log\frac{0.1}{0.1} = 4.74 + 0 = \mathbf{4.74}$$

(Buffer with equal acid & salt concentration $\Rightarrow$ pH = pK_a)

All Concentration Units — Quick Reference

Unit	Formula	Note
Molarity (M)	$\frac{n_{\text{solute}}}{V_{\text{soln}}(\text{L})}$	Temp-dependent
Normality (N)	$\frac{\text{Gram equiv.}}{V(\text{L})}$	$N = M \times n$
Molality (m)	$\frac{n_{\text{solute}}}{m_{\text{solvent}}(\text{kg})}$	Temp-independent
Mole fraction	$\frac{n_A}{n_A + n_B}$	Dimensionless
%(w/w)	$\frac{w_s}{w_{\text{soln}}} \times 100$	Most common

Exam Strategy

Identify quantity asked $\rightarrow$ select formula $\rightarrow$ check units (g/mol, mL/L) before substituting.

Ostwald's Dilution Law

Statement

For a weak electrolyte $HA \rightleftharpoons H^+ + A^-$ at concentration C (mol/L):

$$K_a = \frac{C\alpha^2}{1-\alpha}$$

For weak electrolytes ($\alpha \ll 1$): $K_a \approx C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$

PYQ Numerical (Su'23)

α of CH_3COOH at 0.1 M, $K_a = 1.8 \times 10^{-5}$:

$\alpha = \sqrt{1.8 \times 10^{-5}/0.1} = \sqrt{1.8 \times 10^{-4}} = \mathbf{1.34 \times 10^{-2}}$ (1.34%)

PYQ Discussion: 7-Mark Questions on Unit 1

Frequently Asked (All Papers)

1. Define & explain **buffer solution** with mechanism and buffer equation. (W'25, Su'25)
2. Compare **ionic, covalent, metallic** bonds. (W'23, Su'24)
3. Write electronic configuration of **Cr and Cu**; explain exceptions. (W'23)
4. Explain **Hund's rule** with example; state Pauli's exclusion principle. (Su'24)
5. Define molarity, normality, molality with formulas and numericals. (Su'23, W'25)
6. Explain **degree of ionization**; state factors affecting it. (Su'24, W'24)

Tip

Always write **definition + formula + example + application** for 7-mark answers.

Unit 1 — Key Formulas & Definitions

Topic	Formula / Key Fact
Max e^- in shell	$2n^2$
Molarity	$M = \frac{w \times 1000}{M_w \times V_{mL}}$
Normality	$N = M \times n$-factor
pH	$pH = -\log[H^+]$; pH+pOH=14
H-H Equation	$pH = pK_a + \log \frac{[Salt]}{[Acid]}$
Ostwald's Law	$\alpha = \sqrt{K_a/C}$ (dilute)
APF (FCC/HCP)	74%; BCC: 68%; SC: 52%

Next: Frames 46–60

Advanced topics: Lewis structures, colligative properties, PYQ pattern analysis, Unit 1 mind map

Lewis (Electron Dot) Structures

What is a Lewis Structure?

Representation of molecules/ions showing **all valence electrons** as dots and shared pairs as lines (bonds).

Steps to Draw

1. Count total valence electrons
2. Place least electronegative atom at centre
3. Form single bonds between all atoms
4. Complete octets of outer atoms using remaining e^-
5. If central atom lacks octet, form multiple bonds

PYQ (W'24)

Draw Lewis structures of: H_2O (2 lone pairs on O), NH_3 (1 lone pair on N), CO_2 (double bonds, linear).

Ethene (C_2H_4) – σ and π Bonds (PYQ – W'24)

Structure

$H_2C{=}CH_2$: each C is sp^2 hybridized

Double bond $= 1\sigma + 1\pi$

σ Bond

Head-on overlap

Stronger, shorter

Free rotation

π Bond

Lateral/side overlap

Weaker, longer

No free rotation

Bond Count Rule

Single bond $= 1\sigma$; Double bond $= 1\sigma{+}1\pi$; Triple bond $= 1\sigma{+}2\pi$

NaCl Crystal Structure (Ionic Solid)

Face-Centred Cubic Lattice

- Na^+ and Cl^- alternating in 3D FCC lattice
- Each ion surrounded by **6 oppositely-charged** ions (CN = 6)
- Unit cell contains 4 NaCl formula units
- Held by strong **electrostatic** (ionic) forces

Ionic Solid Properties

High m.p./b.p. · Hard but brittle · Non-conductor (solid)
Conducts in molten/dissolved state · Soluble in polar solvents

Diamond vs Graphite — Network (Covalent) Solids

Diamond

sp^3 hybridized C

3D network, tetrahedral

Hardest natural substance

Non-conductor

m.p. 3550 °C

Use: cutting tools, gems

Graphite

sp^2 hybridized C

Layered hexagonal sheets

Soft, lubricant

Good conductor (delocalized π)

m.p. >3500 °C

Use: electrodes, pencils

Key (PYQ – Su'24)

Both are allotropes of carbon. Diamond = insulator; Graphite = conductor due to delocalized electrons.

Anomalous Properties of Water (H-Bond Effects)

Why Water is Unusual

All anomalous properties of H_2O are due to **extensive H-bonding** between molecules.

Property	**Explanation**
High b.p. (100 °C)	H-bonds require extra energy to break
Ice less dense than water	Hexagonal open lattice in ice $\Rightarrow$ lower density
High surface tension	Strong intermolecular H-bonds
High specific heat capacity	Energy absorbed in breaking H-bonds
Universal solvent	Dipole nature dissolves polar/ionic substances

Quantum Numbers — Summary Table

QN	Symbol	Values	Determines
Principal	n	1, 2, 3, ...	Shell, energy, size
Azimuthal	l	0 to $n-1$	Shape of orbital
Magnetic	m_l	$-l$ to $+l$	Orientation in space
Spin	m_s	$+\frac{1}{2}, -\frac{1}{2}$	Electron spin

Example: 2p electron

$n = 2,\ l = 1,\ m_l \in \{-1, 0, +1\},\ m_s = \pm\frac{1}{2}$

Three $2p$ orbitals: $2p_x$, $2p_y$, $2p_z$ — each holds 2 e^-

Colligative Properties of Solutions

Definition

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

Four Colligative Properties

Property	Formula / Effect
Vapour pressure lowering	$\Delta P = \chi_{\text{solute}} \cdot P^{\circ}$
Boiling point elevation	$\Delta T_b = K_b \cdot m$
Freezing point depression	$\Delta T_f = K_f \cdot m$
Osmotic pressure	$\pi = MRT$

Application

Anti-freeze (ethylene glycol in water): lowers freezing point by $\Delta T_f = K_f \cdot m$

Ionic vs Covalent vs Metallic — Property Comparison

Property	Ionic	Covalent	Metallic
M.p./B.p.	High	Low–High	High
Conductivity	Molten/soln.	Poor	Excellent
Solubility	Polar solvents	Non-polar solvents	Insoluble
Brittleness	Brittle	Varies	Ductile
Bonding	Electrostatic	Shared pairs	e^- sea
Example	NaCl	CH_4, Diamond	Fe, Cu

PYQ Practice: Short-Answer (2-Mark Questions)

Frequently Asked Definitions

Q1 Define molarity and give its SI unit.
Q2 State Hund's rule of maximum multiplicity.
Q3 What is a buffer solution? Give one example.
Q4 Define degree of ionization.
Q5 Write electronic configuration of Cr ($Z = 24$).
Q6 Distinguish between orbit and orbital.
Q7 State Aufbau principle.
Q8 What is coordinate bond? Give example.

Tip

2-mark answers: 2–3 lines max. Use **definition + example** format.

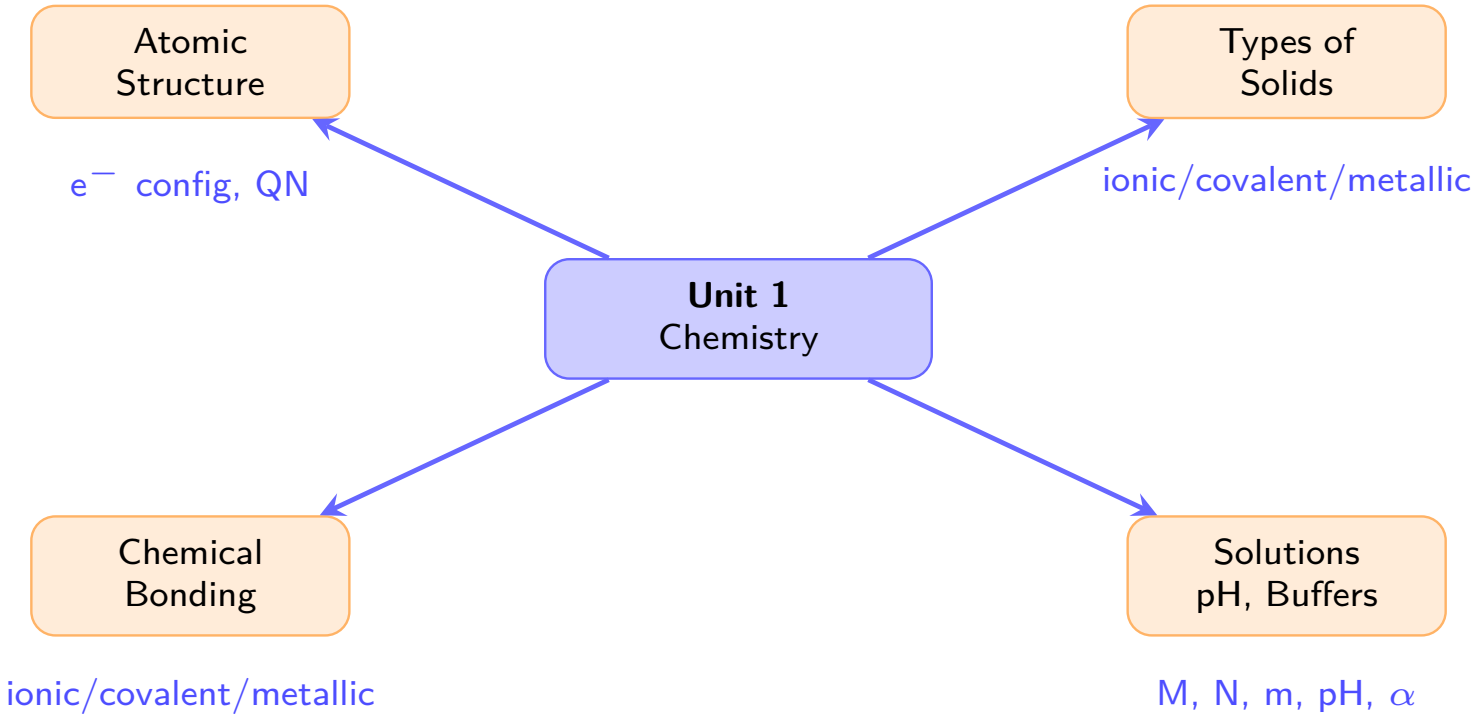
Unit 1 — PYQ Pattern Analysis

Topic	Su23	W23	Su24	W24	Su25	W25
Electronic config.	✓	✓	✓		✓	
Chemical bonding		✓	✓	✓		
Molarity/Normality	✓			✓	✓	
pH calculations		✓		✓		✓
Buffer solutions			✓			✓
Degree of ionization			✓			
Types of solids				✓	✓	

High Probability

Chemical bonding, molarity numericals, and buffer solutions appear in **almost every paper**.

Unit 1 — Mind Map



Unit 1 — Exam Tips

✓ Do's

- Write Cr/Cu config with reason
- Mention units in numericals
- Draw diagrams for bonds
- State all 3 quantum numbers in electronic config questions
- Use H-H equation for buffer pH

✗ Don'ts

- Don't confuse orbit vs orbital
- Don't swap molarity and molality
- Don't forget n-factor in normality
- Don't leave numericals without final box
- Don't mix acidic/basic buffer compositions

Time Allocation

7-mark: $\sim$8 min · 4-mark: $\sim$4 min · 2-mark: $\sim$2 min

Normality Numericals Practice (PYQ – W'25)

Problem 1

Find normality of 6.3 g HNO_3 in 200 mL solution.

$M_w = 63,\ n = 1,\ E_w = 63$

$$N = \frac{6.3 \times 1000}{63 \times 200} = \mathbf{0.5\ N}$$

Problem 2

Find normality of 4 g NaOH in 500 mL.

$M_w = 40,\ n = 1,\ E_w = 40$

$$N = \frac{4 \times 1000}{40 \times 500} = \mathbf{0.2\ N}$$

pH of Weak Acid — Numerical (PYQ – Su'25)

Problem

Find pH of 0.1 M CH_3COOH solution.
$K_a = 1.8 \times 10^{-5}$

Solution (step-by-step)

Step 1: $\alpha = \sqrt{K_a/C} = \sqrt{1.8 \times 10^{-5}/0.1} = 1.34 \times 10^{-2}$

Step 2: $[H^+] = C\alpha = 0.1 \times 1.34 \times 10^{-2} = 1.34 \times 10^{-3}$ M

Step 3: $pH = -\log(1.34 \times 10^{-3}) = 3 - \log(1.34) = 3 - 0.127 = \mathbf{2.87}$

Unit 1 Complete — Preview of Unit 2

Unit 1 Covered

- ✓ Atomic structure, quantum numbers, electronic configuration
- ✓ Ionic, covalent, coordinate, metallic, H-bond, van der Waals
- ✓ Types of solids, crystal structures (FCC/BCC/HCP)
- ✓ Molarity, normality, molality, mole fraction, % concentration
- ✓ pH, pOH, buffer solutions, degree of ionization

Unit 2: Electrochemistry (Coming Next)

Oxidation & Reduction · Electrochemical Cells

Electrode Potential · Nernst Equation

Conductance · Faraday's Laws · Electrolysis Applications