



CHEMISTRY TRICKS & TIPS

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1. MOLE CONCEPT & STOICHIOMETRY

- Moles $n = \text{given mass} / \text{molar mass (g mol}^{-1}\text{)}$.
- Particles = $n \times 6.022 \times 10^{23}$ (Avogadro constant).
- Limiting reagent: compare available moles / balanced-equation coefficient; the smallest ratio limits product.
- Molarity $M = \text{moles of solute} / \text{litres of solution}$.
- Dilution only: $M_1V_1 = M_2V_2$. Keep volume units consistent.
- Empirical formula: convert masses to moles, then divide each by the smallest mole value.

2. ATOMIC STRUCTURE

- $c = \nu \times \lambda$; photon energy $E = h \nu = hc/\lambda$.
- de Broglie wavelength: $\lambda = h/(m v)$.
- Hydrogen-like atom: $E_n = -13.6 Z^2/n^2 \text{ eV}$; $r = 0.529 n^2/Z \text{ \AA}$.
- Maximum electrons in shell $n = 2n^2$; orbitals in shell = n^2 .
- Subshell capacities: $s = 2, p = 6, d = 10, f = 14$ electrons.
- Quantum numbers: $l = 0$ to $n-1$; each orbital holds max 2 electrons with opposite spins.

3. PERIODIC TRENDS

- Across a period: atomic radius generally decreases; ionisation energy and electronegativity generally rise.
- Down a group: atomic radius generally rises; ionisation energy and electronegativity generally fall.
- More effective nuclear charge usually pulls electrons closer; extra shells increase size down a group.
- Isoelectronic species: more protons = smaller radius (e.g., $\text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+}$).
- Common IE exceptions: $\text{Be} > \text{B}$ and $\text{N} > \text{O}$, due to subshell energies and half-filled p stability.
- Use trends as a guide, then check known exceptions and the actual species/charge.

4. BONDING & MOLECULAR SHAPE

- Steric number = sigma bonds + lone pairs on central atom.
- Hybridisation shortcut: SN 2 sp (linear); SN 3 sp^2 (trigonal planar); SN 4 sp^3 (tetrahedral electron geometry).
- Repulsion: $\text{LP-LP} > \text{LP-BP} > \text{BP-BP}$; lone pairs compress bond angles.
- MO bond order = $(\text{bonding } e^- - \text{antibonding } e^-)/2$.
- Higher bond order usually means a shorter, stronger bond.
- Symmetric molecules can have zero net dipole (CO_2); shape matters as well as bond polarity.

5. GASES & GAS LAWS

- Ideal gas: $P V = n R T$; use kelvin and consistent units.
- Gas density $d = P M / (R T)$, where M is molar mass.
- Graham law: $\text{rate}_1/\text{rate}_2 = \sqrt{M_2/M_1}$ for comparable conditions.
- At the same T , lighter gases have higher rms speed: $u_{\text{rms}} = \sqrt{3RT/M}$.
- At 273.15 K and 1 bar, ideal molar volume is about 22.7 L/mol; at 1 atm, about 22.4 L/mol.
- For gas mixtures, total pressure = sum of partial pressures.

6. THERMODYNAMICS

- Heat: $q = m c \Delta T$; phase change: $q = m L$.
- Chemistry sign convention: $\Delta U = q + w$; expansion work $w = -P_{\text{ext}} \Delta V$.
- For ideal-gas reactions: $\Delta H = \Delta U + \Delta n_g R T$.
- Gibbs energy: $\Delta G = \Delta H - T \Delta S$; at constant T, P , $\Delta G < 0$ is spontaneous.
- Hess law: add reaction steps and their enthalpy changes to get the net change.
- Average bond-enthalpy estimate: bonds broken - bonds formed.



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1. CHEMICAL EQUILIBRIUM

- Write K using product concentrations over reactant concentrations, each raised to its balanced coefficient.
- Pure solids and pure liquids are omitted from K expressions.
- For gases:** $K_p = K_c (RT)^{\Delta n_g}$.
- Compare Q with K: $Q < K$ forward; $Q > K$ reverse; $Q = K$ at equilibrium.
- Pressure favours the side with fewer gas moles; a catalyst changes rate, not K or equilibrium position.
- Only temperature changes the equilibrium constant for a given reaction.

2. ACIDS, BASES & BUFFERS

- At 25 C:** $pH = -\log[H^+]$; $pOH = -\log[OH^-]$; $pH + pOH = 14$.
- For a strong monoprotic acid, $[H^+]$ is approximately its concentration; for a strong base, account for OH^- per formula unit.
- Weak acid:** $[H^+] \approx \sqrt{K_a \times C}$ when ionisation is small.
- Buffer:** $pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$.
- A buffer resists pH change; it works best when acid and conjugate-base amounts are comparable.
- Do not use $pH + pOH = 14$ at temperatures other than 25 C unless K_w is specified.

3. SOLUTIONS & COLLIGATIVE PROPERTIES

- Raoult law:** partial vapour pressure $p_A = x_A \times p_A(\text{pure})$.
- Boiling rise:** $\Delta T_b = i K_b m$; **freezing drop:** $\Delta T_f = i K_f m$.
- Osmotic pressure:** $\pi = i C R T$ (C in mol/L with matching R units).
- Colligative properties depend on number of dissolved particles, not their chemical identity.
- Use van't Hoff factor i for dissociation/association; ideal non-electrolyte dilute solution has i approx 1.
- Molality m = moles solute / kg solvent; unlike molarity, it is not temperature-dependent.

4. REDOX QUICK METHODS

- Oxidation = loss of electrons / oxidation number rises; reduction = gain of electrons / oxidation number falls.
- Oxidising agent gets reduced; reducing agent gets oxidised.
- Element in its free form has oxidation number 0; a monatomic ion has its ionic charge.
- O is usually -2, peroxide O is -1; H is usually +1, but metal hydrides have H = -1.
- Sum of oxidation numbers equals the total charge of the species.
- When balancing, separate oxidation and reduction changes; balance atoms and charge, then check atoms and charge again.

5. ELECTROCHEMISTRY

- $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$ (use standard reduction potentials).
- $\Delta G = -n F E_{\text{cell}}$; positive E_{cell} means spontaneous as written under those conditions.
- At 298 K:** $E = E^\circ - (0.0591/n) \log Q$.
- Electrons flow externally from anode to cathode; oxidation occurs at anode, reduction at cathode.
- Conductance** $G = 1/R$; **conductivity** $\kappa = \text{cell constant} / R$.
- During electrolysis, deposited mass is proportional to charge passed: $Q = I t$.

6. CHEMICAL KINETICS

- Rate law:** $\text{rate} = k[A]^m[B]^n$; overall order = m + n from experiment, not automatically from equation coefficients.
- First order:** $k = (2.303/t) \log(a/(a-x))$; $t_{1/2} = 0.693/k$.
- For a first-order reaction, half-life is independent of initial concentration.
- Arrhenius:** $k = A \exp(-E_a/RT)$.
- Higher temperature usually raises rate by increasing the fraction of molecules above activation energy.
- Units of k depend on overall reaction order; always check before selecting an answer.



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1. COORDINATION COMPOUNDS

- Coordination number = number of donor atoms directly bonded to the central metal, not always the number of ligands.
- Name ligands alphabetically; use prefixes di-, tri-, tetra- for counts (these prefixes do not decide alphabetic order).
- Write metal oxidation state in Roman numerals; an anionic complex uses a metal name ending in -ate.
- **Common shapes: CN 2 linear; CN 4 tetrahedral or square planar; CN 6 octahedral.**
- Strong-field ligands can cause pairing; weak-field ligands often give high-spin complexes (consider metal and geometry).
- Isomer checks: ionisation, linkage, coordination, geometrical and optical isomerism.

2. d/f-BLOCK & INORGANIC TRENDS

- Transition metals often show variable oxidation states, coloured ions, complex formation and catalytic activity.
- Unpaired d electrons commonly cause paramagnetism; all paired electrons give diamagnetism.
- d0 and d10 ions are often colourless, but colour exceptions can occur due to charge-transfer transitions.
- Across a period, oxidation states and properties vary; learn important NCERT compounds and their reactions.
- For p-block, focus on oxidation states, anomalous first-element behaviour, oxides, hydrides and key reactions.
- Use short comparison tables for colour, oxidation state, magnetic behaviour and reagent observations.

3. GOC: EFFECTS & STABILITY

- Inductive effect travels through sigma bonds and generally weakens with distance; resonance delocalises electrons.
- Acidity rises when the conjugate base is better stabilised; compare resonance, electronegativity, size and inductive effects.
- Basicity depends on lone-pair availability; resonance delocalisation can make a lone pair less available.
- Carbocation stability: resonance-stabilised benzylic/allylic species are strong; simple alkyl trend is 3 deg > 2 deg > 1 deg.
- For simple alkyl carbanions without resonance, the trend is usually opposite: methyl > 1 deg > 2 deg > 3 deg.
- Hyperconjugation and resonance stabilise intermediates; always compare structures in the actual question.

4. ISOMERISM & HYDROCARBONS

- Structural isomers differ in connectivity; stereoisomers share connectivity but differ in 3D arrangement.
- Maximum stereoisomers for n independent chiral centres is 2^n ; symmetry/meso forms can reduce the number.
- Geometrical isomerism requires restricted rotation and appropriate different groups on each end of a double bond.
- Markovnikov addition usually places H on the carbon already bearing more H atoms.
- Peroxide anti-Markovnikov effect is characteristic of HBr, not generally HCl or HI.
- Terminal alkynes are weakly acidic; benzene typically favours substitution to preserve aromaticity.

5. ORGANIC REACTION CUES

- Aqueous KOH often favours substitution; alcoholic KOH with heat often favours elimination (check substrate).
- Tollens reagent gives a silver mirror with many aldehydes; Fehling solution is positive for many aliphatic aldehydes.
- Aldol reaction: carbonyl compound with an alpha-H; Cannizzaro: aldehyde without an alpha-H under strong base.
- Alcohol oxidation depends on class and reagent; primary alcohols can reach aldehydes/acids, secondary alcohols ketones.
- CN- source matters: KCN tends to give nitrile; AgCN often gives isocyanide as the major product.
- For conversions, write the functional group first, then choose a reagent that performs the required change.

6. SMART NEET CHEMISTRY SOLVING

- Physical: write units, identify the governing formula and check significant powers of ten.
- Inorganic: learn NCERT tables, exceptions, colours, oxidation states and reaction conditions in short revision cycles.
- Organic: identify functional group, electron-rich/electron-poor site, reagent role and likely intermediate.
- Use elimination carefully: reject options that violate charge, atom balance, trend or known reaction conditions.
- After each mock test, sort errors into concept, memory, calculation, reading or time-management errors.
- Revise the error notebook repeatedly; redo missed questions without looking at the solution.