



CBSE Class 11 Chemistry Formula Sheet

67 formulas across 9 chapters — 2025-26 CBSE board exams

Ch 1. Basic Concepts of Chemistry

Number of moles (mol)

$$n = m/M$$

m = mass (g), M = molar mass (g/mol)

Number of particles

$$N = n \times N_a$$

$N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$ (Avogadro's number)

Molarity (mol/L)

$$M = n_{\text{solute}} / V_{\text{solution(L)}}$$

Mole fraction

$$\chi_A = n_A / (n_A + n_B + \dots)$$

Parts per million

$$\text{ppm} = (\text{mass of solute} / \text{mass of solution}) \times 10^6$$

Percentage yield

$$\% \text{ yield} = (\text{actual yield} / \text{theoretical yield}) \times 100$$

Tip: Theoretical yield is based on limiting reagent

Limiting reagent identification

Divide moles of each reactant by its stoichiometric coefficient — smallest ratio → limiting reagent

Empirical formula steps

% → g (assume 100g) → mol ($\div M$) → simple ratio → empirical formula

Tip: Multiply ratios to get whole numbers if needed

Ch 2. Structure of Atom

Bohr's radius (hydrogen)

$$r_n = n^2 \times 0.529 \text{ \AA}$$

n = principal quantum number

Tip: For hydrogen-like ions: $r_n = (n^2/Z) \times 0.529 \text{ \AA}$

Energy of nth Bohr orbit

$$E_n = -13.6 / n^2 \text{ eV}$$

Tip: Negative sign indicates bound state; energy increases (less negative) as n increases

Rydberg formula

$$1/\lambda = R_H (1/n_1^2 - 1/n_2^2)$$

$R_H = 1.097 \times 10^7 \text{ m}^{-1}$, $n_1 < n_2$

Tip: $n_1=1$: Lyman (UV), $n_1=2$: Balmer (visible), $n_1=3$: Paschen (IR)

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de Broglie wavelength

$$\lambda = h / mv = h / p$$

$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$, m = mass, v = velocity

Heisenberg uncertainty principle

$$\Delta x \cdot \Delta p \geq h / 4\pi$$

Tip: Δx = uncertainty in position, Δp = uncertainty in momentum; cannot simultaneously know both precisely

Max electrons in a shell

$$2n^2$$

n = principal quantum number ($K=2$, $L=8$, $M=18$, $N=32$)

Max electrons in a subshell

$$2(2l + 1)$$

l = azimuthal quantum number ($s:2$, $p:6$, $d:10$, $f:14$)

Number of orbitals in subshell

$$(2l + 1)$$

Tip: s : 1 orbital, p : 3, d : 5, f : 7

Ch 3. Classification of Elements and Periodicity

Effective nuclear charge

$$Z_{\text{eff}} = Z - \sigma$$

Z = atomic number, σ = shielding constant (from Slater's rules)

Tip: Electrons farther from nucleus experience lower Z_{eff} due to shielding

Electronegativity (Mulliken scale)

$$EN = (IE_1 + EA) / 2$$

IE_1 = first ionisation enthalpy, EA = electron affinity

Tip: Pauling scale is more common; F has highest $EN = 4.0$

Atomic radius trend (period)

Decreases left $\rightarrow$ right ($\uparrow Z_{\text{eff}}$ pulls electrons closer)

Atomic radius trend (group)

Increases top $\rightarrow$ bottom (new shells added)

Ionisation enthalpy trend (period)

Increases left $\rightarrow$ right (with exceptions at Group 2 $\rightarrow$ 3 and 5 $\rightarrow$ 6)

Tip: $IE_1 < IE_2 < IE_3 \dots$ always; large jump in IE indicates valence shell is exhausted

Electron gain enthalpy trend

Most negative for Group 17 (halogens); Cl more negative than F due to small size of F

Tip: More negative = more energy released = higher tendency to gain electron

Metallic character trend

Decreases across period, increases down group

Ch 4. Chemical Bonding and Molecular Structure

Bond order (MO theory)

$$BO = (N_b - N_a) / 2$$

N_b = bonding electrons, N_a = antibonding electrons

Tip: Higher bond order $\rightarrow$ shorter bond length, higher bond enthalpy, more stable molecule

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Formal charge

$$FC = V - N - B/2$$

V = valence electrons, N = non-bonding electrons, B = shared (bonding) electrons

Tip: Structures with FC closest to zero on each atom are most stable

Dipole moment

$$\mu = q \times d$$

q = charge (in esu), d = bond length (cm); 1 Debye = 3.336×10^{-30} C·m

Tip: Net dipole = vector sum of all bond dipoles; symmetrical molecules (BF_3 , CCl_4) have $\mu = 0$

VSEPR notation

$$\text{AX}_n\text{E}_m$$

A = central atom, X = bonded atoms (n), E = lone pairs (m)

Tip: AX_4E_0 = tetrahedral; AX_3E_1 = pyramidal; AX_2E_2 = bent (V-shape)

Hybridisation and geometry

sp : linear | sp^2 : trigonal planar | sp^3 : tetrahedral | sp^3d : trigonal bipyramidal | sp^3d^2 : octahedral

Bond length order

Triple bond < Double bond < Single bond

Tip: More electrons shared → atoms pulled closer → shorter bond

Resonance: stability

More resonance structures → greater delocalisation → more stable molecule

Tip: Resonance hybrid has lower energy than any single contributing structure

Ch 5. Thermodynamics

Relation between ΔH and ΔU

$$\Delta H = \Delta U + \Delta n_g RT$$

Δn_g = moles of gaseous products – moles of gaseous reactants

Tip: For reactions with no gas: $\Delta H \approx \Delta U$

Hess's law

$$\Delta H_{\text{reaction}} = \sum \Delta H_{\text{products}} - \sum \Delta H_{\text{reactants}} \quad (\text{using standard enthalpies})$$

Tip: Enthalpy is a state function — path doesn't matter

Bond enthalpy method

$$\Delta H_{\text{rxn}} = \sum \text{BE}(\text{bonds broken}) - \sum \text{BE}(\text{bonds formed})$$

Tip: Energy is absorbed to break bonds and released to form bonds

Gibbs free energy

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

T = temperature in K, ΔS = entropy change

Spontaneity condition

$\Delta G < 0 \rightarrow$ spontaneous | $\Delta G = 0 \rightarrow$ equilibrium | $\Delta G > 0 \rightarrow$ non-spontaneous

ΔG and equilibrium constant

$$\Delta G^\circ = -RT \ln K$$

Tip: $K > 1 \rightarrow \Delta G^\circ < 0 \rightarrow$ products favoured

First law of thermodynamics

$$\Delta U = Q + W \quad (\text{IUPAC sign convention})$$

Tip: Note: $W = -P\Delta V$; work done ON system is positive

Ch 6. Equilibrium

Equilibrium constant K_c

$$K_c = \frac{[\text{products}]^{\text{stoich}}}{[\text{reactants}]^{\text{stoich}}} \text{ (at equilibrium)}$$

Equilibrium constant K_p

$$K_p = K_c(RT)^{\Delta n_g}$$

$$R = 0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K}, T \text{ in K}$$

Ionic product of water

$$K_w = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

Acid dissociation constant

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Base dissociation constant

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

Relation $K_a \times K_b$

$$K_a \times K_b = K_w \text{ (conjugate acid-base pair)}$$

Tip: Stronger acid has weaker conjugate base

pH definition

$$\text{pH} = -\log[\text{H}^+] \quad | \quad \text{pOH} = -\log[\text{OH}^-] \quad | \quad \text{pH} + \text{pOH} = 14$$

Henderson-Hasselbalch equation

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

Tip: Used to find pH of buffer solutions

Degree of dissociation (weak acid)

$$\alpha = \sqrt{K_a/C}$$

C = initial concentration of weak acid

Tip: Valid when $\alpha \ll 1$

pH of weak acid

$$\text{pH} = \frac{1}{2}(\text{p}K_a - \log C)$$

Tip: Easier form of Henderson equation for pure weak acid

Ch 7. Redox Reactions

Oxidation number: free element

$$\text{ON} = 0 \text{ (e.g. Fe, Cl}_2, \text{S}_8)$$

Oxidation number: monoatomic ion

$$\text{ON} = \text{charge of ion (e.g. Na}^+ \rightarrow +1, \text{Cl}^- \rightarrow -1)$$

Oxidation number: oxygen

$$\text{ON(O)} = -2 \text{ usually; } -1 \text{ in peroxides (H}_2\text{O}_2); +2 \text{ in OF}_2$$

Oxidation number: hydrogen

$$\text{ON(H)} = +1 \text{ usually; } -1 \text{ in metal hydrides (NaH)}$$

n-factor (acid-base)

$$\text{n-factor} = \text{basicity of acid or acidity of base}$$

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n-factor (redox)

n-factor = change in oxidation number per formula unit

Tip: In ionic equations: n-factor = total electrons transferred per formula unit

Equivalents (for calculations)

Equivalents = moles $\times$ n-factor | $N = M \times$ n-factor

Half-reaction method (acidic)

Add H_2O to balance O, then H^+ to balance H, then e^- to balance charge

Ch 8. Organic Chemistry — Basic Principles and Techniques

Index of Hydrogen Deficiency (IHD)

$$IHD = (2C + 2 + N - H - X) / 2$$

C = carbons, N = nitrogens, H = hydrogens, X = halogens; O and S do not affect IHD

Tip: IHD = 1: one ring or one double bond; IHD = 2: one triple bond or two degrees of unsaturation

Inductive effect (+I order)



Tip: +I groups donate electrons; -I groups (halogens, NO_2 , CN) withdraw electrons

Carbocation stability



Tip: Hyperconjugation and inductive effect of alkyl groups stabilise higher-order carbocations

Carbanion stability



Tip: Electron-withdrawing groups stabilise carbanions; alkyl groups destabilise them

Free radical stability



Tip: Hyperconjugation stabilises free radicals; used in halogenation of alkanes

IUPAC priority for functional groups



Tip: Highest-priority group gives suffix; others become prefixes

Ch 9. Hydrocarbons

Markovnikov's rule

In addition to unsymmetrical alkene: H adds to C with more H atoms; X adds to C with fewer H atoms

Tip: Explains regioselectivity; anti-Markovnikov occurs with HBr in presence of peroxides (radical mechanism)

Hückel's rule (aromaticity)

Aromatic if: planar, fully conjugated ring, $(4n + 2)$ π electrons ($n = 0, 1, 2, \dots$)

Tip: Benzene: $n=1$, 6π electrons; naphthalene: 10π ; cyclopentadienyl anion: 6π

Free radical halogenation selectivity

Reactivity: $F_2 > Cl_2 > Br_2 > I_2$ | Selectivity: $I_2 > Br_2 > Cl_2 > F_2$

Tip: Br_2 is most useful: high selectivity for $3^\circ > 2^\circ > 1^\circ$ H atoms

Baeyer's reagent test

Alkene/Alkyne + cold dil. $KMnO_4$ (alk.) $\rightarrow$ colourless (purple decolourises) = unsaturation confirmed

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Ozonolysis

$C=C + O_3 \rightarrow$ [ozonide] $\rightarrow$ with Zn/H_2O : two carbonyl fragments | with H_2O_2 : carboxylic acids

Tip: Used to locate double bond position by identifying carbonyl fragments

EAS — directing effects

ortho/para directors (activating): $-OH$, $-NH_2$, $-CH_3$, halogens | meta directors (deactivating): $-NO_2$

Tip: Halogens are deactivating yet o/p directors — remember as an exception