CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 1 : Some Basic Concepts of Chemistry (NCERT)

1.1 Laws of Chemical Combination

Law	Statement
Law of Conservation of Mass	Matter can neither be created nor destroyed (Lavoisier). Total mass of reactants = total mass of products.
Law of Definite Proportions	A compound always contains the same elements in the same fixed ratio by mass (Proust).
Law of Multiple Proportions	If two elements form more than one compound, masses of one element combining with a fixed mass of the other are in a simple whole-number ratio (Dalton). e.g. H_2O & H_2O_2 → O ratio 1 : 2
Gay-Lussac's Law of Gaseous Volumes	Gases combine in simple whole-number ratios by volume at the same T & P.
Avogadro's Law	Equal volumes of all gases at the same T & P contain equal number of molecules.

1.2 Mole Concept — Master Formulas

Quantity	Formula / Expression
Number of moles (n)	n = Given mass (w) ÷ Molar mass (M)
From particles	n = Number of particles ÷ N_A ($N_A = 6.022 \times 10^{23}$ mol^{-1})
From gas volume	n = Volume at STP ÷ 22.4 L (at 1 atm, 273 K) | ÷ 22.7 L (at 1 bar)
Number of particles	Particles = n × N_A ; Atoms = molecules × atomicity
Mass of 1 molecule	= Molar mass ÷ N_A (in grams)
Vapour density (VD)	Molar mass = 2 × Vapour density
Average atomic mass	= Σ (fractional abundance × isotopic mass)
1 amu (u)	= 1/12 mass of one C-12 atom = 1.66×10^{-24} g

1.3 Percentage Composition, Empirical & Molecular Formula

Concept	Formula / Expression
Mass % of element	= (n × Atomic mass of element ÷ Molar mass of compound) × 100
Empirical formula (steps)	1) % ÷ atomic mass → moles 2) divide by smallest 3) make whole numbers
Molecular formula	MF = n × EF, where n = Molar mass ÷ Empirical formula mass
Example	EF = CH_2O, M = 180 g/mol → n = 180/30 = 6 → MF = $C_6H_{12}O_6$ (glucose)

1.4 Concentration Terms

Term	Formula	Unit / Note
Mass percent	(Mass of solute ÷ Mass of solution) × 100	%, T-independent
Mole fraction (x_A)	$x_A = n_A \div (n_A + n_B)$	$x_A + x_B = 1$, no unit

Term	Formula	Unit / Note
Molarity (M)	Moles of solute ÷ Volume of solution (L)	mol/L, changes with T
Molality (m)	Moles of solute ÷ Mass of solvent (kg)	mol/kg, T-independent
ppm	(Mass of solute ÷ Mass of solution) × 10^6	for trace amounts

1.5 Dilution & Mixing of Solutions

Situation	Formula
Dilution (same solute)	$M_1V_1 = M_2V_2$
Mixing two solutions (same solute)	$M_{mix} = (M_1V_1 + M_2V_2) \div (V_1 + V_2)$
Molarity from mass % (d = density g/mL)	M = (10 × mass % × d) ÷ Molar mass
Neutralisation / titration	$M_1V_1 \times n_1 = M_2V_2 \times n_2$ (n = basicity/acidity)

1.6 Stoichiometry & Limiting Reagent

Concept	How to Apply
Stoichiometric calculations	Balance equation → convert given data to moles → use mole ratio → convert to required quantity (mass/volume/particles).
Limiting reagent (LR)	Divide moles of each reactant by its coefficient; the SMALLEST value = limiting reagent. LR decides the amount of product.
Excess reagent left	Moles left = initial moles − (moles consumed as per LR ratio)
Percentage yield	% yield = (Actual yield ÷ Theoretical yield) × 100
Example	3 mol H_2 + 1 mol N_2 : $N_2 + 3H_2 \rightarrow 2NH_3$; H_2: 3/3 = 1, N_2: 1/1 = 1 → exact ratio, no LR.

1.7 Uncertainty in Measurement — Significant Figures

Rule	Example
All non-zero digits are significant	285 cm → 3 sig. figs.
Zeros between non-zero digits are significant	5005 → 4 sig. figs.
Leading zeros are NOT significant	0.0025 → 2 sig. figs.
Trailing zeros after decimal ARE significant	0.200 → 3 sig. figs.
Addition/Subtraction	Result keeps least DECIMAL PLACES (12.11 + 18.0 = 30.1)
Multiplication/Division	Result keeps least SIG. FIGS. (2.5 × 1.25 = 3.1)
Scientific notation	$N \times 10^n$, where $1 \leq N < 10$ (e.g. 232.508 = 2.32508×10^2)

1.8 Key Values to Memorise

Quantity	Value
Avogadro number (N_A)	6.022×10^{23} mol^{-1}

Quantity	Value
Molar volume of gas	22.4 L at STP (1 atm, 273.15 K) ; 22.7 L at 1 bar, 273.15 K
Molar masses (g/mol)	H_2O = 18, CO_2 = 44, NH_3 = 17, CH_4 = 16, O_2 = 32, N_2 = 28, NaCl = 58.5, H_2SO_4 = 98, $CaCO_3$ = 100, NaOH = 40, HCl = 36.5, $KMnO_4$ = 158
1 L	= 1 dm^3 = 1000 mL = 1000 cm^3

★ QUICK MEMORY AIDS & TRICKS

Mole bridge	GRAMS ↔ MOLES ↔ PARTICLES ↔ LITRES — moles sit at the centre; always convert to moles first!
Molarity vs Molality	Molar–Litre (both have L sound) ; Molal–kilogram of solvent. Molality never changes with temperature.
LR shortcut	moles ÷ coefficient → smallest wins (is limiting).
M from %	M = 10·x·d / M_{solute} — remember '10xd' as 'ten times density'.
22.4 vs 22.7	Old STP (1 atm) → 22.4 L ; New NCERT STP (1 bar) → 22.7 L.
Sig. fig. zeros	'Sandwich zeros count, leading zeros never, trailing zeros only after a decimal.'

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 1 : Some Basic Concepts of Chemistry (NCERT)

1.1 Laws of Chemical Combination

Law	Statement
Law of Conservation of Mass	Matter can neither be created nor destroyed (Lavoisier). Total mass of reactants = total mass of products.
Law of Definite Proportions	A compound always contains the same elements in the same fixed ratio by mass (Proust).
Law of Multiple Proportions	If two elements form more than one compound, masses of one element combining with a fixed mass of the other are in a simple whole-number ratio (Dalton). e.g. H_2O & H_2O_2 → O ratio 1 : 2
Gay-Lussac's Law of Gaseous Volumes	Gases combine in simple whole-number ratios by volume at the same T & P.
Avogadro's Law	Equal volumes of all gases at the same T & P contain equal number of molecules.

1.2 Mole Concept — Master Formulas

Quantity	Formula / Expression
Number of moles (n)	n = Given mass (w) ÷ Molar mass (M)
From particles	n = Number of particles ÷ N_A ($N_A = 6.022 \times 10^{23}$ mol^{-1})
From gas volume	n = Volume at STP ÷ 22.4 L (at 1 atm, 273 K) | ÷ 22.7 L (at 1 bar)
Number of particles	Particles = n × N_A ; Atoms = molecules × atomicity
Mass of 1 molecule	= Molar mass ÷ N_A (in grams)
Vapour density (VD)	Molar mass = 2 × Vapour density
Average atomic mass	= Σ (fractional abundance × isotopic mass)
1 amu (u)	= 1/12 mass of one C-12 atom = 1.66×10^{-24} g

1.3 Percentage Composition, Empirical & Molecular Formula

Concept	Formula / Expression
Mass % of element	= (n × Atomic mass of element ÷ Molar mass of compound) × 100
Empirical formula (steps)	1) % ÷ atomic mass → moles 2) divide by smallest 3) make whole numbers
Molecular formula	MF = n × EF, where n = Molar mass ÷ Empirical formula mass
Example	EF = CH_2O, M = 180 g/mol → n = 180/30 = 6 → MF = $C_6H_{12}O_6$ (glucose)

1.4 Concentration Terms

Term	Formula	Unit / Note
Mass percent	(Mass of solute ÷ Mass of solution) × 100	%, T-independent
Mole fraction (x_A)	$x_A = n_A \div (n_A + n_B)$	$x_A + x_B = 1$, no unit

Term	Formula	Unit / Note
Molarity (M)	Moles of solute ÷ Volume of solution (L)	mol/L, changes with T
Molality (m)	Moles of solute ÷ Mass of solvent (kg)	mol/kg, T-independent
ppm	(Mass of solute ÷ Mass of solution) × 10^6	for trace amounts

1.5 Dilution & Mixing of Solutions

Situation	Formula
Dilution (same solute)	$M_1V_1 = M_2V_2$
Mixing two solutions (same solute)	$M_{mix} = (M_1V_1 + M_2V_2) \div (V_1 + V_2)$
Molarity from mass % (d = density g/mL)	M = (10 × mass % × d) ÷ Molar mass
Neutralisation / titration	$M_1V_1 \times n_1 = M_2V_2 \times n_2$ (n = basicity/acidity)

1.6 Stoichiometry & Limiting Reagent

Concept	How to Apply
Stoichiometric calculations	Balance equation → convert given data to moles → use mole ratio → convert to required quantity (mass/volume/particles).
Limiting reagent (LR)	Divide moles of each reactant by its coefficient; the SMALLEST value = limiting reagent. LR decides the amount of product.
Excess reagent left	Moles left = initial moles − (moles consumed as per LR ratio)
Percentage yield	% yield = (Actual yield ÷ Theoretical yield) × 100
Example	3 mol H_2 + 1 mol N_2 : $N_2 + 3H_2 \rightarrow 2NH_3$; H_2: 3/3 = 1, N_2: 1/1 = 1 → exact ratio, no LR.

1.7 Uncertainty in Measurement — Significant Figures

Rule	Example
All non-zero digits are significant	285 cm → 3 sig. figs.
Zeros between non-zero digits are significant	5005 → 4 sig. figs.
Leading zeros are NOT significant	0.0025 → 2 sig. figs.
Trailing zeros after decimal ARE significant	0.200 → 3 sig. figs.
Addition/Subtraction	Result keeps least DECIMAL PLACES (12.11 + 18.0 = 30.1)
Multiplication/Division	Result keeps least SIG. FIGS. (2.5 × 1.25 = 3.1)
Scientific notation	$N \times 10^n$, where $1 \leq N < 10$ (e.g. 232.508 = 2.32508×10^2)

1.8 Key Values to Memorise

Quantity	Value
Avogadro number (N_A)	6.022×10^{23} mol^{-1}

Quantity	Value
Molar volume of gas	22.4 L at STP (1 atm, 273.15 K) ; 22.7 L at 1 bar, 273.15 K
Molar masses (g/mol)	H_2O = 18, CO_2 = 44, NH_3 = 17, CH_4 = 16, O_2 = 32, N_2 = 28, NaCl = 58.5, H_2SO_4 = 98, $CaCO_3$ = 100, NaOH = 40, HCl = 36.5, $KMnO_4$ = 158
1 L	= 1 dm^3 = 1000 mL = 1000 cm^3

★ QUICK MEMORY AIDS & TRICKS

Mole bridge	GRAMS ↔ MOLES ↔ PARTICLES ↔ LITRES — moles sit at the centre; always convert to moles first!
Molarity vs Molality	Molar–Litre (both have L sound) ; Molal–kilogram of solvent. Molality never changes with temperature.
LR shortcut	moles ÷ coefficient → smallest wins (is limiting).
M from %	M = 10·x·d / M_{solute} — remember '10xd' as 'ten times density'.
22.4 vs 22.7	Old STP (1 atm) → 22.4 L ; New NCERT STP (1 bar) → 22.7 L.
Sig. fig. zeros	'Sandwich zeros count, leading zeros never, trailing zeros only after a decimal.'

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 2 : Structure of Atom (NCERT)

2.1 Subatomic Particles & Atomic Terms

Particle / Term	Key Values / Definition
Electron (Thomson)	Charge = -1.602×10^{-19} C ; mass = 9.1×10^{-31} kg ; e/m = 1.758×10^{11} C/kg
Proton	Charge = $+1.602 \times 10^{-19}$ C ; mass = 1.672×10^{-27} kg
Neutron (Chadwick)	Charge = 0 ; mass = 1.675×10^{-27} kg
Atomic number (Z)	Z = number of protons = number of electrons (neutral atom)
Mass number (A)	A = protons + neutrons → Neutrons = A − Z
Isotopes	Same Z, different A (e.g. ^{1}H, ^{2}H, ^{3}H)
Isobars	Same A, different Z (e.g. ^{14}C and ^{14}N)
Isotones	Same number of neutrons (e.g. ^{14}C and ^{16}O — 8 neutrons each)

2.2 Electromagnetic Radiation & Planck's Theory

Quantity	Formula / Expression
Speed–frequency–wavelength	$c = \nu \times \lambda$ ($c = 3 \times 10^{8}$ m/s)
Wave number ($\bar{\nu}$)	$\bar{\nu} = 1/\lambda$ (unit: m^{-1} or cm^{-1})
Energy of one photon	$E = h\nu = hc/\lambda$ ($h = 6.626 \times 10^{-34}$ J·s)
Energy of n photons	$E = nh\nu$
Handy shortcut	$E\ (eV) = 12400 \div \lambda\ (Å)$ or $E\ (J) = 1.986 \times 10^{-25} \div \lambda\ (m)$
Photoelectric effect (Einstein)	$h\nu = h\nu_0 + \frac{1}{2}mv^2 \rightarrow KE_{max} = h\nu - W_0$ (W_0 = work function = $h\nu_0$)
Threshold condition	Electron ejected only if $\nu \geq \nu_0$ (higher intensity → more electrons, NOT higher energy)

2.3 Bohr's Model (H and H-like species: He^{+}, Li^{2+})

Quantity	Formula	Quick Values
Radius of nth orbit	$r_n = 52.9 \times n^2/Z$ pm	$r_1(H) = 52.9$ pm $= 0.529$ Å
Energy of nth orbit	$E_n = -13.6 \times Z^2/n^2$ eV $= -2.18 \times 10^{-18} \times Z^2/n^2$ J	$E_1(H) = -13.6$ eV
Velocity of electron	$v_n = 2.18 \times 10^{6} \times Z/n$ m/s	$v_1(H) \approx c/137$
Angular momentum	$mvr = nh/2\pi$ (quantised)	n = 1, 2, 3 …
Ionisation energy	$IE = E_\infty - E_1 = +13.6\ Z^2$ eV	H: 13.6 eV; He^{+}: 54.4 eV
Energy gap	$\Delta E = E_2 - E_1 = 13.6\ Z^2\ (1/n_1^2 - 1/n_2^2)$ eV	= hν of emitted photon

2.4 Hydrogen Spectrum — Rydberg Formula

Concept	Formula / Detail
Rydberg formula	$1/\lambda = R_H \times Z^2\ (1/n_1^2 - 1/n_2^2)$; $R_H = 109677\ cm^{-1}$

Concept	Formula / Detail
Lyman series	$n_1 = 1$, $n_2 = 2,3,4\ldots \rightarrow$ Ultraviolet region
Balmer series	$n_1 = 2$, $n_2 = 3,4,5\ldots \rightarrow$ VISIBLE region (only visible series)
Paschen series	$n_1 = 3 \rightarrow$ Infrared
Brackett series	$n_1 = 4 \rightarrow$ Infrared
Pfund series	$n_1 = 5 \rightarrow$ Infrared
Max. spectral lines	From level n: total lines = $n(n-1)/2$; between n_2 and n_1: $(n_2-n_1)(n_2-n_1+1)/2$

2.5 Dual Nature & Heisenberg Uncertainty Principle

Concept	Formula / Expression
de Broglie wavelength	$\lambda = h/mv = h/p$
With kinetic energy	$\lambda = h \div \sqrt{(2m \cdot KE)}$
Charged particle (charge q, accelerated by V volts)	$\lambda = h \div \sqrt{(2mqV)}$; for electron: $\lambda = 12.27/\sqrt{V}$ Å
Heisenberg uncertainty principle	$\Delta x \cdot \Delta p \geq h/4\pi$ or $\Delta x \cdot \Delta v \geq h/(4\pi m)$
Significance	Meaningful only for microscopic particles; rules out fixed Bohr orbits → 'orbitals' (probability regions)

2.6 Quantum Numbers

Quantum Number	Symbol & Values	What It Tells
Principal	n = 1, 2, 3, … (K, L, M, N…)	Shell, size & energy of orbital
Azimuthal (angular)	l = 0 to (n−1) ; s=0, p=1, d=2, f=3	Subshell & shape of orbital
Magnetic	$m_l = -l \ldots 0 \ldots +l$ (total 2l+1 values)	Orientation of orbital
Spin	$m_s = +½$ or $-½$	Spin of electron (↑ or ↓)

Counting Formula	Expression
Orbitals in a shell	n^2 ; Max electrons in shell = $2n^2$
Orbitals in a subshell	2l + 1 ; Max electrons in subshell = 2(2l + 1) → s:2, p:6, d:10, f:14
Radial nodes	n − l − 1
Angular nodes	l ; Total nodes = n − 1
Orbital angular momentum	$= \sqrt{(l(l+1))} \times h/2\pi$
Spin angular momentum	$= \sqrt{(s(s+1))} \times h/2\pi$, s = ½
Spin magnetic moment	$\mu = \sqrt{(n(n+2))}$ BM (n = number of unpaired electrons)

2.7 Rules for Filling Electrons

Rule	Statement
Aufbau principle	Electrons fill orbitals in order of increasing energy: lower (n + l) first; if (n + l) equal → lower n first
Energy order	1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s
Pauli exclusion principle	No two electrons in an atom can have all four quantum numbers identical (max 2 e^- per orbital, opposite spins)
Hund's rule	Pairing starts only after each orbital of a subshell gets one electron (maximum multiplicity)
Exceptional configs	Cr (Z=24): $[Ar]3d^5 4s^1$; Cu (Z=29): $[Ar]3d^{10} 4s^1$ — extra stability of half/fully-filled subshells
Cation formation	Remove electrons from HIGHEST n first (4s before 3d): $Fe^{2+} = [Ar]3d^6$

★ QUICK MEMORY AIDS & TRICKS

Spectrum series order	'Lyman Bhai Pass Bhi Phir-se' → Lyman, Balmer, Paschen, Brackett, Pfund (n_1 = 1,2,3,4,5)
−13.6 eV table	E_n(H): n=1 → −13.6 ; n=2 → −3.4 ; n=3 → −1.51 ; n=4 → −0.85 eV (divide by n^2)
Lines shortcut	Drop from n=4 to 1 → 4(3)/2 = 6 lines.
Nodes	Radial = n − l − 1, Angular = l, Total = n − 1 → '3p: 1 radial + 1 angular = 2 total'.
de Broglie quick	Heavier or faster particle → smaller λ (why cricket balls don't show waves!)
12400 rule	E(eV) × λ(Å) = 12400 — fastest way for photon numericals.
Orbital capacity	s-2, p-6, d-10, f-14 → 'Spdf: 2, 6, 10, 14 — add 4 each time'.

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 3 : Classification of Elements and Periodicity in Properties (NCERT)

3.1 Development of the Periodic Table

Scientist / Law	Key Idea
Dobereiner's Triads	Groups of 3 elements; atomic mass of middle element ≈ average of other two (e.g. Li-7, Na-23, K-39).
Newlands' Law of Octaves	Every 8th element has properties similar to the 1st (worked only up to Ca).
Mendeleev's Periodic Law	Properties are a periodic function of **atomic masses**. Left gaps and predicted Ga (eka-aluminium) and Ge (eka-silicon).
Modern Periodic Law (Moseley)	Properties are a periodic function of **atomic number (Z)**. Basis: $\sqrt{v} = a(Z - b)$.

3.2 Structure of the Modern Periodic Table

Feature	Details
Periods (7)	Elements per period: 1st → 2, 2nd → 8, 3rd → 8, 4th → 18, 5th → 18, 6th → 32, 7th → 32
Groups (18)	Vertical columns; same valence-shell configuration → similar properties
s-block (Gr 1, 2)	ns^{1-2} ; alkali & alkaline earth metals; highly reactive
p-block (Gr 13-18)	ns^2np^{1-6} ; s + p = Representative elements; Gr 18 = noble gases (ns^2np^6)
d-block (Gr 3-12)	$(n-1)d^{1-10}ns^{0-2}$; transition metals; coloured ions, variable oxidation states
f-block	$(n-2)f^{1-14}(n-1)d^{0-1}ns^2$; Lanthanoids (4f) + Actinoids (5f) = inner transition
Metalloids	B, Si, Ge, As, Sb, Te — border the zig-zag line between metals and non-metals
Group number trick	For ns^anp^b elements: Group = 10 + a + b ; for ns^a: Group = a

3.3 IUPAC Nomenclature of Elements with Z > 100

Digit	Root	Digit	Root
0	nil (n)	5	pent (p)
1	un (u)	6	hex (h)
2	bi (b)	7	sept (s)
3	tri (t)	8	oct (o)
4	quad (q)	9	enn (e)

Example	Name & Symbol
Z = 101	Unnilunium (Unu) → official: Mendelevium (Md)
Z = 104	Unnilquadium (Unq) → official: Rutherfordium (Rf)
Z = 120	Unbinilium (Ubn) — add roots in order of digits + 'ium'

3.4 Effective Nuclear Charge & Shielding

Concept	Formula / Statement
Effective nuclear charge	$Z_{eff} = Z - \sigma$ (σ = shielding/screening constant)
Shielding power of subshells	s > p > d > f (s-electrons shield best)
Across a period	Z_{eff} INCREASES (electrons added to same shell shield poorly)
Down a group	Z_{eff} nearly constant; new shells dominate → size increases

3.5 Atomic & Ionic Radius

Concept	Rule / Trend
Types of radii	Covalent radius < Metallic radius < van der Waals radius
Across a period	Radius DECREASES (Z_{eff} ↑) — Li > Be > B > C > N > O > F
Down a group	Radius INCREASES (new shell added) — Li < Na < K < Rb < Cs
Cation vs parent atom	Cation is SMALLER (fewer e^-, higher Z_{eff}): $Na^+ < Na$
Anion vs parent atom	Anion is LARGER (more e^-–e^- repulsion): $Cl^- > Cl$
Isoelectronic series	Same number of electrons → MORE protons = SMALLER size: $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

3.6 Ionization Enthalpy (ΔiH)

Concept	Rule / Trend
Definition	Energy required to remove the outermost electron from an isolated gaseous atom: $X(g) \rightarrow X^+(g) + e^-$
Successive IE	$\Delta_iH_1 < \Delta_iH_2 < \Delta_iH_3$ … (each removal from a more positive ion)
Across a period	Generally INCREASES (Z_{eff} ↑, size ↓)
Down a group	DECREASES (size ↑, shielding ↑)
Exception: Be > B	Be ($2s^2$ fully filled) needs more energy than B (removes easy $2p^1$)
Exception: N > O	N ($2p^3$ half-filled, extra stable) > O ($2p^4$, pairing repulsion)
Factors	↑ with Z_{eff} & half/fully-filled stability ; ↓ with size & shielding

3.7 Electron Gain Enthalpy (ΔegH)

Concept	Rule / Trend
Definition	Enthalpy change when an isolated gaseous atom gains an electron: $X(g) + e^- \rightarrow X^-(g)$
Sign	Usually NEGATIVE (energy released); 2nd $\Delta_{eg}H$ is always POSITIVE ($O^- + e^- \rightarrow O^{2-}$ needs energy)
Across a period	More negative (halogens most negative in their period)
Down a group	Less negative (size ↑)
Exception: Cl > F	Cl has the MOST negative $\Delta_{eg}H$ (F is too small — strong e^-–e^- repulsion in compact 2p)
Noble gases	Positive $\Delta_{eg}H$ (stable ns^2np^6 configuration resists extra electron)

3.8 Electronegativity (EN)

Concept	Rule / Trend
Definition	Tendency of an atom in a COMPOUND to attract shared electrons (not measurable for isolated atom; no units)
Scales	Pauling scale (most common): F = 4.0 (highest), O = 3.5, N = Cl = 3.0; Cs ≈ 0.7 (lowest)
Across a period	INCREASES (Li 1.0 → F 4.0)
Down a group	DECREASES (F > Cl > Br > I)
Relation	EN ∝ Z_{eff} ∝ 1/size ; high EN → non-metallic character

3.9 Master Trend Table

Property	Across Period (→)	Down Group (↓)
Atomic radius	Decreases	Increases
Ionization enthalpy	Increases	Decreases
Electron gain enthalpy (magnitude)	Increases	Decreases
Electronegativity	Increases	Decreases
Metallic character	Decreases	Increases
Non-metallic character	Increases	Decreases
Z_{eff}	Increases	Almost constant
Basic nature of oxides	Decreases (basic → amphoteric → acidic)	Increases

3.10 Valence, Oxides & Anomalous Behaviour

Concept	Rule / Example
Valence (groups 1-14)	= number of valence electrons; (groups 15-18) = 8 − valence electrons
Oxides across period 3	Na_2O (basic) → MgO (basic) → Al_2O_3 (amphoteric) → SiO_2 (weakly acidic) → P_2O_5, SO_3, Cl_2O_7 (acidic)
Anomalous 2nd period	Li, Be, B, C differ from their groups: small size, high EN, no d-orbitals, max covalence = 4
Diagonal relationship	Li ↔ Mg, Be ↔ Al, B ↔ Si (similar size & EN diagonally)
Reactivity pattern	Highest at the two extremes (Gr 1 metals & Gr 17 halogens); lowest in the middle

★ QUICK MEMORY AIDS & TRICKS

Z > 100 roots	'nil un bi tri quad pent hex sept oct enn' → 0-9; just read digits of Z and add 'ium'.
IE exceptions	'Be beats B, N beats O' — full $2s^2$ and half-filled $2p^3$ are extra stable.
$\Delta_{eg}H$ champion	Cl (not F!) has most negative electron gain enthalpy: Cl > F > Br > I.
Isoelectronic size	More protons pull harder → smaller ion. 'Cations shrink, anions swell.'
EN order to memorise	F (4.0) > O (3.5) > Cl = N (3.0) > Br (2.8) > S = C (2.5) > H (2.1)
Group number trick	p-block: Group = 10 + (s electrons) + (p electrons). e.g. $3s^23p^5$ → 10+2+5 = Group 17 (Cl).

Metalloid mnemonic	'B Si Ge As Sb Te' → 'Basi GeAs SabTe' — the staircase elements.

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 4 : Chemical Bonding and Molecular Structure (NCERT)

4.1 Lewis Structures, Octet Rule & Formal Charge

Concept	Formula / Rule
Octet rule	Atoms bond to attain 8 electrons in the valence shell (noble-gas configuration)
Formal charge (FC)	FC = V − L − ½B (V = valence e^- of free atom, L = lone-pair e^-, B = bonding/shared e^-)
Best Lewis structure	The one with formal charges closest to ZERO; negative FC on the more electronegative atom
Incomplete octet	LiCl, BeH_2, BCl_3, BF_3 (central atom has < 8 e^-)
Expanded octet	PF_5 (10 e^-), SF_6 (12 e^-), H_2SO_4, IF_7 (only 3rd period onward — d-orbitals)
Odd-electron molecules	NO (11 e^-), NO_2 (17 e^-) — octet rule fails
Resonance	Real structure = hybrid of all canonical forms; more resonance → more stability (O_3, CO_3^{2-}, benzene)

4.2 Ionic Bond, Lattice Enthalpy & Fajan's Rules

Concept	Rule / Statement
Ionic bond formation favoured by	Low ionization enthalpy of metal + high (negative) electron gain enthalpy of non-metal + high lattice enthalpy
Lattice enthalpy	Energy to completely separate 1 mole of ionic solid into gaseous ions; ↑ with charge, ↓ with ionic radius
Fajan's rules (covalent character ↑ if)	Cation: small & highly charged ; Anion: large & highly charged ; cation with pseudo noble-gas config (e.g. Cu^+)
Polarising power order	$Li^+ > Na^+ > K^+$ (smaller cation polarises more) ; covalent character: LiCl > NaCl > KCl

4.3 Bond Parameters & Dipole Moment

Property	Formula / Relation
Bond order (VBT)	= number of bonds between two atoms (C−C: 1, C=C: 2, C≡C: 3)
Bond order relations	Bond order ↑ → Bond enthalpy ↑ & Bond length ↓
Isoelectronic species	Same bond order: N_2, CO, NO^+, CN^- (all BO = 3)
Dipole moment (μ)	μ = q × d ; unit: Debye (D), 1 D = 3.336 × 10^{-30} C·m
% Ionic character	= (Observed μ ÷ Theoretical μ for 100% ionic) × 100
Zero dipole (symmetric)	CO_2, BF_3, CH_4, CCl_4, PCl_5, SF_6 — bond dipoles cancel
Non-zero dipole	H_2O (1.85 D), NH_3 (1.47 D) > NF_3 (0.23 D — lone pair opposes N−F dipoles)
Dipole order trick	cis-isomer has μ > 0; trans (symmetric) often μ = 0

4.4 VSEPR Theory — Geometry from BP + LP

BP + LP	BP	LP	Shape	Angle	Example
2	2	0	Linear	180°	$BeCl_2$, CO_2
3	3	0	Trigonal planar	120°	BF_3
3	2	1	Bent / V-shape	< 120°	SO_2, $SnCl_2$
4	4	0	Tetrahedral	109.5°	CH_4, NH_4^+
4	3	1	Trigonal pyramidal	107°	NH_3
4	2	2	Bent	104.5°	H_2O
5	5	0	Trigonal bipyramidal	120°, 90°	PCl_5
5	4	1	See-saw	—	SF_4
5	3	2	T-shape	—	ClF_3
5	2	3	Linear	180°	XeF_2, I_3^-
6	6	0	Octahedral	90°	SF_6
6	5	1	Square pyramidal	—	BrF_5
6	4	2	Square planar	90°	XeF_4

VSEPR Rule	Statement
Repulsion order	Lone pair–Lone pair > Lone pair–Bond pair > Bond pair–Bond pair
Effect of lone pairs	Each lone pair squeezes bond angle: CH_4 (109.5°) → NH_3 (107°) → H_2O (104.5°)

4.5 Valence Bond Theory & Hybridisation

Concept	Formula / Rule
Hybridisation shortcut	H = ½ (V + M − C + A) ; V = valence e^- of central atom, M = monovalent atoms, C = cationic charge, A = anionic charge
H value → type	2 → sp ; 3 → sp^2 ; 4 → sp^3 ; 5 → sp^3d ; 6 → sp^3d^2
Alternative count	Hybridisation = σ-bonds + lone pairs on central atom
Sigma vs pi	σ: head-on overlap (stronger, free rotation) ; π: sideways overlap (weaker); double bond = 1σ + 1π, triple = 1σ + 2π
% s-character & angle	sp (50% s, 180°) > sp^2 (33%, 120°) > sp^3 (25%, 109.5°); more s-character → shorter, stronger bond
Bond strength of overlap	s–s < s–p < p–p (for σ overlap of same period)

Hybridisation	Geometry	Bond Angle	Examples
sp	Linear	180°	$BeCl_2$, CO_2, C_2H_2
sp^2	Trigonal planar	120°	BF_3, C_2H_4, SO_3
sp^3	Tetrahedral	109.5°	CH_4, CCl_4, NH_4^+
sp^3d	Trigonal bipyramidal	120° & 90°	PCl_5
sp^3d^2	Octahedral	90°	SF_6
d^2sp^3 / sp^3d^2	Octahedral	90°	$[Co(NH_3)_6]^{3+}$ (Class 12 preview)

4.6 Molecular Orbital Theory (MOT)

Concept	Formula / Rule
Bond order (MOT)	BO = ½ (N_b − N_a) ; N_b = bonding e^-, N_a = antibonding e^-
Stability	BO > 0 → molecule exists ; BO = 0 → does NOT exist (He_2, Be_2)
Magnetism	Unpaired e^- → paramagnetic ; all paired → diamagnetic
Energy order (up to N_2, Z ≤ 14)	$\sigma1s < \sigma^*1s < \sigma2s < \sigma^*2s < (\pi2p_x = \pi2p_y) < \sigma2p_z < (\pi^*2p_x = \pi^*2p_y) < \sigma^*2p_z$
Energy order (O_2, F_2)	$\sigma2p_z$ comes BEFORE $\pi2p_x = \pi2p_y$
Adding to antibonding MO	Removes ½ from BO → O_2^+ (BO 2.5) > O_2 (2) > O_2^- (1.5) > O_2^{2-} (1)

Species	Total e^-	Bond Order	Magnetism
H_2	2	1	Diamagnetic
He_2	4	0 (does not exist)	—
Li_2	6	1	Diamagnetic
B_2	10	1	Paramagnetic (2 unpaired)
C_2	12	2	Diamagnetic
N_2	14	3	Diamagnetic
O_2	16	2	Paramagnetic (2 unpaired)
F_2	18	1	Diamagnetic
Ne_2	20	0 (does not exist)	—

4.7 Hydrogen Bonding

Concept	Rule / Example
Condition	H attached to highly electronegative F, O, or N
Strength order	H…F > H…O > H…N (but weaker than covalent bond, about 10-40 kJ/mol)
Intermolecular H-bond	Between molecules: H_2O, HF, alcohols → high b.p., water's anomalous density
Intramolecular H-bond	Within one molecule: o-nitrophenol, salicylaldehyde → LOWER b.p. than para-isomer
Why ice floats	Open cage structure due to H-bonding → ice less dense than liquid water

★ QUICK MEMORY AIDS & TRICKS

FC in seconds	FC = V − L − B/2 → 'Valence minus Lone minus half-Bonded'.
Hybridisation trick	H = ½(V + M − C + A); SF_4: ½(6+4) = 5 → sp^3d ; NH_4^+: ½(5+4−1) = 4 → sp^3.
BO ladder for O_2 family	O_2^+ 2.5 > O_2 2 > O_2^- 1.5 > O_2^{2-} 1 — each added e^- goes to π* and cuts 0.5.
Paramagnetic pair	'B_2 and O_2 are the odd ones out' — both paramagnetic with 2 unpaired electrons.
Angle squeeze	109.5° → 107° → 104.5° : each lone pair eats 2.5° (CH_4 → NH_3 → H_2O).
NH_3 vs NF_3 dipole	NH_3 (1.47 D) >> NF_3 (0.23 D): lone pair adds in NH_3 but opposes in NF_3.
Zero-dipole list	'CO_2, BF_3, CH_4, CCl_4, PCl_5, SF_6' — perfectly symmetric = μ = 0.

Isoelectronic BO = 3 club	N_2, CO, NO^+, CN^- — all 14 e^-, all bond order 3.

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 5 : Thermodynamics (NCERT)

5.1 Basic Terms & System Types

Term	Definition / Detail
System	Part of the universe under study
Surroundings	Everything outside the system
Open system	Exchanges both matter & energy (e.g. open beaker)
Closed system	Exchanges only energy, not matter (e.g. sealed flask)
Isolated system	Exchanges neither matter nor energy (e.g. thermos flask)
State function	Depends only on initial & final state: U, H, S, G, T, P, V
Path function	Depends on path taken: heat (q) and work (w)
Extensive property	Depends on amount: mass, volume, U, H, S, G, heat capacity
Intensive property	Independent of amount: T, P, density, molar volume, refractive index

5.2 First Law of Thermodynamics

Concept	Formula / Expression
First law (energy conservation)	$\Delta U = q + w$
Sign convention	Heat absorbed BY system: $q = +$; Work done ON system: $w = +$
Pressure–volume work	$w = -P_{ext}\,\Delta V$ (irreversible) ; $\Delta V = V_{final} - V_{initial}$
Reversible isothermal (ideal gas)	$w = -2.303\ nRT \log(V_2/V_1) = -2.303\ nRT \log(P_1/P_2)$
Free expansion (into vacuum)	$P_{ext} = 0 \rightarrow w = 0$
Isothermal ideal gas	$\Delta U = 0$ and $\Delta T = 0 \rightarrow q = -w$
Adiabatic process	$q = 0 \rightarrow \Delta U = w$ (temperature changes)
Isochoric (constant V)	$\Delta V = 0 \rightarrow w = 0$, so $q_v = \Delta U$
Isobaric (constant P)	$q_p = \Delta H$

5.3 Enthalpy (H) & Heat Capacity

Concept	Formula / Expression
Enthalpy definition	$H = U + PV$; $\Delta H = \Delta U + P\Delta V$
For gaseous reactions	$\Delta H = \Delta U + \Delta n_g RT$; Δn_g = moles of gaseous products − gaseous reactants
Heat at constant V	$q_v = \Delta U$; Heat at constant P: $q_p = \Delta H$
Specific heat	$q = m \times c \times \Delta T$ (c = specific heat capacity, J g^{-1} K^{-1})
Molar heat capacity	$q = n \times C \times \Delta T$ (C = molar heat capacity, J mol^{-1} K^{-1})
Relation (ideal gas)	$C_p - C_v = R$ (per mole)
Heat capacity ratio	$\gamma = C_p/C_v$ (monatomic 1.67, diatomic 1.40)

Concept	Formula / Expression
C_v values	Monatomic: (3/2)R ; Diatomic: (5/2)R

5.4 Enthalpies of Reactions (Thermochemistry)

Type of Enthalpy	Definition
Enthalpy of reaction ($\Delta_r H$)	Heat change when reaction occurs in molar quantities
Standard conditions	Standard state: 1 bar pressure, specified T (usually 298 K)
Enthalpy of formation ($\Delta_f H°$)	Heat change forming 1 mol compound from elements in standard states (elements: $\Delta_f H° = 0$)
Enthalpy of combustion ($\Delta_c H°$)	Heat released on complete combustion of 1 mol substance (always negative)
Enthalpy of fusion ($\Delta_{fus} H°$)	Heat to melt 1 mol solid → liquid at m.p.
Enthalpy of vaporisation ($\Delta_{vap} H°$)	Heat to vaporise 1 mol liquid → gas at b.p.
Enthalpy of sublimation ($\Delta_{sub} H°$)	$\Delta_{sub}H = \Delta_{fus}H + \Delta_{vap}H$
Enthalpy of atomisation ($\Delta_a H°$)	Heat to break 1 mol substance into gaseous atoms
Bond enthalpy	Energy to break 1 mol of a particular bond in gaseous state
Lattice / Solution / Dilution	$\Delta_{lattice}H$, $\Delta_{sol}H$, $\Delta_{dil}H$ for ionic solids & solutions

5.5 Hess's Law & Enthalpy Calculations

Concept	Formula
Hess's law of constant heat summation	Total ΔH is the same whether reaction occurs in one step or several steps (since H is a state function)
From formation enthalpies	$\Delta_r H° = \Sigma\, \Delta_f H°(\text{products}) - \Sigma\, \Delta_f H°(\text{reactants})$
From bond enthalpies	$\Delta_r H = \Sigma$ B.E.(bonds broken, reactants) − Σ B.E.(bonds formed, products)
From combustion enthalpies	$\Delta_r H° = \Sigma\, \Delta_c H°(\text{reactants}) - \Sigma\, \Delta_c H°(\text{products})$

5.6 Spontaneity, Entropy & Gibbs Energy

Concept	Formula / Expression
Entropy (S)	Measure of disorder/randomness; a state function. $\Delta S = q_{rev}/T$
Entropy of surroundings	$\Delta S_{surr} = -\Delta H_{sys}/T$
Second law (total entropy)	For a spontaneous process: $\Delta S_{total} = \Delta S_{sys} + \Delta S_{surr} > 0$
At phase equilibrium	$\Delta S = \Delta H/T$ (e.g. melting: $\Delta S_{fus} = \Delta_{fus}H / T_m$)
Gibbs energy (definition)	$G = H - TS$; $\Delta G = \Delta H - T\Delta S$ (at constant T, P)
Gibbs–Helmholtz criterion	$\Delta G < 0$ → spontaneous ; $\Delta G = 0$ → equilibrium ; $\Delta G > 0$ → non-spontaneous
ΔG and equilibrium constant	$\Delta G° = -RT \ln K = -2.303\, RT \log K$
ΔG and useful work	$-\Delta G$ = maximum non-expansion (useful) work obtainable
Third law of thermodynamics	Entropy of a perfectly crystalline substance is zero at 0 K

5.7 Spontaneity from Signs of ΔH and ΔS

ΔH	ΔS	ΔG = ΔH − TΔS	Spontaneity
− (exo)	+	Always −	Spontaneous at ALL temperatures
+ (endo)	−	Always +	Non-spontaneous at ALL temperatures
− (exo)	−	− at low T	Spontaneous only at LOW temperature
+ (endo)	+	− at high T	Spontaneous only at HIGH temperature

★ QUICK MEMORY AIDS & TRICKS

First law signs	'Heat IN is +, Work ON is +' → ΔU = q + w (chemistry convention).
ΔH vs ΔU	$\Delta H = \Delta U + \Delta n_g RT$ — only GASEOUS moles count in Δn_g. If $\Delta n_g = 0$, ΔH = ΔU.
State vs path	State functions = 'destination only' (U,H,S,G,T,P,V). Path functions = q and w.
Cp − Cv = R	For ideal gas only. Cp is always greater (extra energy for expansion work).
Spontaneity grid	'−ΔH and +ΔS = always spontaneous; +ΔH and −ΔS = never.'
ΔG = ΔH − TΔS	Negative ΔG drives reactions. 'Low T favours enthalpy, high T favours entropy.'
ΔG° = −2.303 RT log K	K > 1 → ΔG° negative (products favoured) ; K < 1 → ΔG° positive.
Δsub = Δfus + Δvap	Solid → gas in two imaginary steps (Hess's law shortcut).

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 6 : Equilibrium (NCERT)

6.1 Equilibrium Constants (Kc & Kp)

Concept	Formula / Expression
For aA + bB ⇌ cC + dD	$K_c = [C]^c[D]^d \div [A]^a[B]^b$ (concentrations at equilibrium)
K_p (gaseous)	$K_p = (P_C^c \times P_D^d) \div (P_A^a \times P_B^b)$
Relation Kp ↔ Kc	$K_p = K_c (RT)^{\Delta n}$; Δn = (moles gaseous products − moles gaseous reactants)
When Δn = 0	$K_p = K_c$ (equal gas moles on both sides)
Units	K is dimensionless when Δn = 0; otherwise has units
Reaction reversed	K' = 1/K
Reaction × n	$K' = K^n$
Reactions added	$K = K_1 \times K_2$ (multiply individual constants)
Pure solids & liquids	Taken as 1 (excluded from K expression)

6.2 Reaction Quotient & Equilibrium Predictions

Concept	Rule
Reaction quotient (Q)	Same expression as K but with non-equilibrium concentrations
Q < K	Forward reaction proceeds (more products form)
Q > K	Reverse reaction proceeds (more reactants form)
Q = K	System is at equilibrium (no net change)
Magnitude of K	$K > 10^3$ → products dominate ; $K < 10^{-3}$ → reactants dominate ; in between → comparable
Link to Gibbs energy	$\Delta G^\circ = -RT \ln K = -2.303\ RT \log K$; $\Delta G^\circ < 0 \rightarrow K > 1$

6.3 Le Chatelier's Principle

Disturbance	Equilibrium Shift
Add reactant	Shifts FORWARD
Add product	Shifts BACKWARD
Increase pressure (↓ volume)	Shifts toward side with FEWER gas moles
Decrease pressure (↑ volume)	Shifts toward side with MORE gas moles
Increase temperature	Shifts in ENDOTHERMIC direction (K changes)
Decrease temperature	Shifts in EXOTHERMIC direction (K changes)
Add inert gas at constant V	NO shift (partial pressures unchanged)
Add inert gas at constant P	Shifts toward side with MORE gas moles
Add catalyst	NO shift; only reaches equilibrium FASTER (K unchanged)

6.4 Acids & Bases — Concepts

Concept	Definition
Arrhenius	Acid gives H^+ in water ; Base gives OH^- in water
Bronsted–Lowry	Acid = proton (H^+) donor ; Base = proton acceptor
Conjugate pair	Acid − H^+ → conjugate base ; differ by one proton (e.g. HCl/Cl^-)
Lewis	Acid = electron-pair acceptor ; Base = electron-pair donor
Strong vs weak	Strong acids/bases ionise completely; weak ones partially (have K_a/K_b)
Strong acids (memorise)	HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$

6.5 Ionisation of Water, pH Scale

Concept	Formula
Ionic product of water	$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 298 K
pH definition	$pH = -\log[H^+]$; $pOH = -\log[OH^-]$
pH + pOH	pH + pOH = 14 (at 298 K only)
Neutral / acidic / basic	pH = 7 neutral ; pH < 7 acidic ; pH > 7 basic (at 298 K)
$[H^+]$ from pH	$[H^+] = 10^{-pH}$
pKw	$pK_w = pH + pOH = 14$

6.6 Weak Acids/Bases — Ostwald's Dilution Law

Concept	Formula
Acid ionisation constant	$K_a = C\alpha^2/(1-\alpha) \approx C\alpha^2$ (when $\alpha << 1$)
Degree of dissociation	$\alpha = \sqrt{(K_a/C)}$
$[H^+]$ of weak acid	$[H^+] = C\alpha = \sqrt{(K_a \times C)}$
$[OH^-]$ of weak base	$[OH^-] = \sqrt{(K_b \times C)}$
pKa, pKb	$pK_a = -\log K_a$; $pK_b = -\log K_b$
Conjugate pair relation	$K_a \times K_b = K_w \rightarrow pK_a + pK_b = 14$
Stronger acid	Larger K_a / smaller pK_a

6.7 Buffers & Salt Hydrolysis

Type	Formula
Acidic buffer (Henderson)	$pH = pK_a + \log([Salt]/[Acid])$
Basic buffer (Henderson)	$pOH = pK_b + \log([Salt]/[Base])$
Salt of weak acid + strong base	Basic solution: $pH = 7 + \frac{1}{2}(pK_a + \log C)$

Type	Formula
Salt of strong acid + weak base	Acidic solution: pH = 7 − ½(pK$_b$ + log C)
Salt of weak acid + weak base	pH = 7 + ½(pK$_a$ − pK$_b$)
Salt of strong acid + strong base	Neutral (pH = 7), no hydrolysis (e.g. NaCl)

6.8 Solubility Product (Ksp)

Salt Type	Ksp Expression (S = molar solubility)
General A_xB_y	$K_{sp} = (xS)^x(yS)^y = x^x\, y^y\, S^{(x+y)}$
AB type (AgCl, $BaSO_4$)	$K_{sp} = S^2 \rightarrow S = \sqrt{K_{sp}}$
AB_2 or A_2B (CaF_2, Ag_2CrO_4)	$K_{sp} = 4S^3$
AB_3 or A_3B	$K_{sp} = 27S^4$
A_2B_3 (Bi_2S_3)	$K_{sp} = 108S^5$
Precipitation condition	Ionic product (Q) > K_{sp} → precipitate forms ; Q = K_{sp} → saturated ; Q < K_{sp} → unsaturated
Common ion effect	Adding a common ion DECREASES solubility (shifts equilibrium back)

★ QUICK MEMORY AIDS & TRICKS

6 strong acids	HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$ → 'Have Brought In Nice Sulphuric Perchloric'.
Kp = Kc when	Δn = 0 (equal gas moles both sides). Otherwise multiply by $(RT)^{\Delta n}$.
Q vs K	'Q Less → goes Forward, Q More → reverses.' (Q < K forward, Q > K backward)
Ka × Kb = Kw	→ pK$_a$ + pK$_b$ = 14. Stronger acid → weaker conjugate base.
Ksp shortcuts	AB → S^2 ; AB_2/A_2B → $4S^3$; AB_3 → $27S^4$; A_2B_3 → $108S^5$.
Le Chatelier inert gas	Constant V → no shift ; constant P → shifts to MORE moles side.
Salt pH rule	Weak acid salt → basic ; weak base salt → acidic ; both strong → neutral.
Buffer best when	pH = pK$_a$ (equal salt & acid → max buffer capacity).

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 7 : Redox Reactions (NCERT)

7.1 Core Concepts — Oxidation & Reduction

Concept	Definition
Oxidation	Loss of electrons / gain of oxygen / loss of hydrogen / INCREASE in oxidation number
Reduction	Gain of electrons / loss of oxygen / gain of hydrogen / DECREASE in oxidation number
Oxidising agent (oxidant)	Accepts electrons; gets REDUCED itself
Reducing agent (reductant)	Donates electrons; gets OXIDISED itself
Redox reaction	Oxidation and reduction occur simultaneously
Memory: OIL RIG	Oxidation Is Loss, Reduction Is Gain (of electrons)

7.2 Rules for Assigning Oxidation Number (ON)

Rule	Detail
Free element	ON = 0 (e.g. H_2, O_2, Na, S_8, P_4)
Monatomic ion	ON = charge on the ion (Na^+ = +1, Cl^- = −1)
Hydrogen	+1 usually; −1 in metal hydrides (NaH, CaH_2)
Oxygen	−2 usually; −1 in peroxides (H_2O_2); −½ in superoxides (KO_2); +2 in OF_2
Fluorine	Always −1 (most electronegative)
Alkali metals (Gr 1)	Always +1
Alkaline earth metals (Gr 2)	Always +2
Sum rule (neutral molecule)	Sum of all oxidation numbers = 0
Sum rule (ion)	Sum of all oxidation numbers = charge on ion

7.3 Types of Redox Reactions

Type	Description & Example
Combination	Two species combine; at least one element changes ON: $C + O_2 \rightarrow CO_2$
Decomposition	Compound breaks into simpler substances: $2H_2O \rightarrow 2H_2 + O_2$
Displacement	One element displaces another: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
Metal displacement	More reactive metal displaces less reactive from salt
Non-metal displacement	$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$
Disproportionation	SAME element simultaneously oxidised AND reduced: $2H_2O_2 \rightarrow 2H_2O + O_2$; $Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$
Comproportionation	Two species of same element form a single intermediate ON product

7.4 Balancing Redox Reactions

Method	Steps
Oxidation number method	1) Assign ON 2) Identify oxidised & reduced atoms 3) Balance change in ON by suitable coefficients 4) Balance remaining atoms & charge
Ion-electron (half-reaction) method	1) Split into oxidation & reduction half-reactions 2) Balance atoms except O & H 3) Balance O with H_2O, H with H^+ 4) Balance charge with e^- 5) Equalise electrons & add
In basic medium	After acidic balancing, add OH^- equal to H^+ on both sides; combine $H^+ + OH^- \rightarrow H_2O$

7.5 Equivalent Concept (Redox Titrations)

Concept	Formula
n-factor (redox)	= total change in oxidation number per formula unit (electrons gained/lost)
Equivalent mass	= Molar mass ÷ n-factor
Equivalents	Number of equivalents = mass ÷ equivalent mass = moles × n-factor
Normality (N)	N = Molarity × n-factor ; N = equivalents ÷ volume (L)
Titration equivalence	$N_1V_1 = N_2V_2$ (equivalents of oxidant = equivalents of reductant)
n-factor of $KMnO_4$ (acidic)	5 (Mn: +7 → +2)
n-factor of $KMnO_4$ (neutral/basic)	3 (Mn: +7 → +4)
n-factor of $K_2Cr_2O_7$ (acidic)	6 (2Cr: +6 → +3 each)

7.6 Redox & Electrode Potential (Intro)

Concept	Detail
Electrochemical cell	Galvanic cell converts chemical energy → electrical energy (spontaneous)
Anode / Cathode	Oxidation at ANODE (−) ; Reduction at CATHODE (+)
Cell EMF	$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
Spontaneity	$E^\circ_{cell} > 0$ → reaction spontaneous ($\Delta G^\circ = -nFE^\circ_{cell}$)
Standard hydrogen electrode (SHE)	Reference electrode, E° = 0.00 V by convention
Reactivity from E°	More negative E° → stronger reducing agent (e.g. Li, K) ; more positive E° → stronger oxidising agent (e.g. F_2)

★ QUICK MEMORY AIDS & TRICKS

OIL RIG	Oxidation Is Loss, Reduction Is Gain (of electrons) — the classic redox mnemonic.
Agent swap	Oxidising agent gets reduced ; Reducing agent gets oxidised (they do the opposite of their name).
Oxygen exceptions	Normal −2 ; peroxide −1 ; superoxide −½ ; OF_2 = +2 (F beats O).
Disproportionation	ONE element goes both up and down. Classic: H_2O_2, Cl_2 in cold alkali, P_4 in NaOH.
$KMnO_4$ n-factor	Acidic = 5, Neutral = 3, Basic = 1 (Mn goes +7 → +2 / +4 / +6).
$Cr_2O_7^{2-}$ n-factor	Always 6 in acidic medium (2 Cr each +6 → +3).

Normality bridge	N = M × n-factor. Titration: $N_1V_1 = N_2V_2$.
Cell electrodes	'An Ox, Red Cat' → ANode OXidation, REDuction CAThode.

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 8 : Organic Chemistry — Some Basic Principles & Techniques (NCERT)

8.1 Degree of Unsaturation (DoU / IHD)

Compound Type	Formula
Only C & H (C_nH_m)	DoU = (2n + 2 − m) ÷ 2
With nitrogen ($C_nH_mN_p$)	DoU = (2n + 2 + p − m) ÷ 2 (each N adds 1)
With halogen X ($C_nH_mX_x$)	DoU = (2n + 2 − m − x) ÷ 2 (halogen counts like H)
With O or S	DoU = (2n + 2 − m) ÷ 2 (O and S have NO effect)
Meaning of DoU	Each double bond OR ring = 1 ; each triple bond = 2
Benzene ring (C_6H_6)	DoU = 4 (3 double bonds + 1 ring)

8.2 IUPAC Nomenclature — Priority of Functional Groups

Priority (high → low)	Group & Suffix
1 (highest)	Carboxylic acid −COOH (-oic acid)
2	Sulphonic acid $-SO_3H$
3	Ester −COOR (-oate)
4	Acid halide −COX (-oyl halide)
5	Amide $-CONH_2$ (-amide)
6	Nitrile −CN (-nitrile)
7	Aldehyde −CHO (-al)
8	Ketone >C=O (-one)
9	Alcohol −OH (-ol)
10	Amine $-NH_2$ (-amine)
Lowest	Ethers, alkenes (-ene), alkynes (-yne); halo/alkyl/nitro are prefixes only

8.3 Quantitative Analysis — Estimation of Elements

Element / Method	Formula (% of element)
Carbon (combustion)	% C = (12/44) × (mass of CO_2 ÷ mass of compound) × 100
Hydrogen (combustion)	% H = (2/18) × (mass of H_2O ÷ mass of compound) × 100
Nitrogen (Dumas)	% N = (28/22400) × (volume of N_2 at STP ÷ mass of compound) × 100
Nitrogen (Kjeldahl)	% N = (1.4 × M_{acid} × basicity × V_{acid}) ÷ mass of compound
Halogen (Carius) — Cl	% Cl = (35.5/143.5) × (mass AgCl ÷ mass compound) × 100
Halogen (Carius) — Br	% Br = (80/188) × (mass AgBr ÷ mass compound) × 100
Halogen (Carius) — I	% I = (127/235) × (mass AgI ÷ mass compound) × 100
Sulphur (Carius)	% S = (32/233) × (mass $BaSO_4$ ÷ mass compound) × 100

Element / Method	Formula (% of element)
Phosphorus (Carius)	% P = (62/222) × (mass $Mg_2P_2O_7$ ÷ mass compound) × 100

8.4 Electronic Effects

Effect	Description & Examples
Inductive effect (−I)	Electron-withdrawing through σ-bonds: $-NO_2$ > −CN > −COOH > −F > −Cl > −Br > −I > −OH
Inductive effect (+I)	Electron-donating: $-C(CH_3)_3$ > $-CH(CH_3)_2$ > $-C_2H_5$ > $-CH_3$ > −H
Effect of −I / +I on acidity	−I groups INCREASE acid strength; +I groups DECREASE it
Mesomeric / Resonance (+M)	Electron-donating via lone pair: −OH, −OR, $-NH_2$, −NHR, $-NR_2$, halogens
Mesomeric / Resonance (−M)	Electron-withdrawing via π-bond: $-NO_2$, −CHO, −COR, −COOH, −COOR, −CN
Hyperconjugation	Delocalisation of σ(C−H) electrons into adjacent empty p / π orbital; more α-H → more stable
Electromeric effect (E)	Temporary, complete shift of π-electrons in presence of an attacking reagent

8.5 Stability of Reaction Intermediates

Intermediate	Stability Order (most → least)
Carbocation (C^+)	3° > 2° > 1° > CH_3^+ (stabilised by +I & hyperconjugation; benzylic/allylic by resonance)
Carbanion (C^-)	CH_3^- > 1° > 2° > 3° (opposite of carbocation; stabilised by −I groups)
Free radical	3° > 2° > 1° > CH_3• (like carbocation; hyperconjugation)
Resonance-stabilised	Allyl & benzyl cations/radicals are extra stable due to resonance

8.6 Bond Fission & Types of Reagents

Concept	Detail
Homolytic fission	Bond breaks evenly → free radicals (each atom keeps 1 e^-); shown by single-barbed arrows
Heterolytic fission	Bond breaks unevenly → ions (cation + anion); shown by double-barbed arrows
Electrophile (E^+)	Electron-deficient; accepts electron pair (e.g. H^+, NO_2^+, Cl^+, carbocation)
Nucleophile (Nu^-)	Electron-rich; donates electron pair (e.g. OH^-, CN^-, NH_3, H_2O)

8.7 Types of Organic Reactions & Isomerism

Type	Description / Example
Substitution	One atom/group replaces another: $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
Addition	Atoms add across multiple bond: $CH_2{=}CH_2 + H_2 \rightarrow CH_3CH_3$
Elimination	Atoms removed; new π-bond forms: $C_2H_5OH \rightarrow CH_2{=}CH_2 + H_2O$
Rearrangement	Skeleton reorganises (e.g. 1° carbocation → 2°/3°)
Structural isomerism	Chain, position, functional, metamerism, tautomerism

Type	Description / Example
Stereoisomerism	Geometrical (cis-trans) & optical (mirror images)

8.8 Purification Techniques

Method	Used For
Crystallisation	Solids with different solubility from impurities
Sublimation	Solids that sublime (camphor, naphthalene)
Simple distillation	Liquids with large boiling-point difference
Fractional distillation	Liquids with close boiling points (e.g. petroleum fractions)
Steam distillation	Steam-volatile, water-immiscible liquids (e.g. aniline)
Distillation under reduced pressure	Liquids that decompose at their boiling point (e.g. glycerol)
Differential extraction	Separates compound from aqueous solution using a solvent
Chromatography	Separation & purification based on differential adsorption/partition

★ QUICK MEMORY AIDS & TRICKS

DoU shortcut	(2C + 2 + N − H − X) ÷ 2. O & S ignored. Benzene = 4 (3 π + 1 ring).
Priority order	'COOH beats everything' → acid > ester > amide > nitrile > aldehyde > ketone > alcohol > amine.
Carbocation stability	3° > 2° > 1° > methyl. Carbanion is the EXACT reverse.
−I acidity	More −I (electron-withdrawing) groups → stronger acid (stabilises the conjugate base).
E vs Nu	Electrophile = Electron-loving (+) ; Nucleophile = Nucleus-loving / electron-rich (−).
Estimation fractions	C: 12/44, H: 2/18, S: 32/233, halogen via Ag salt mass ratios.
Hyperconjugation	'No-bond resonance' — more α-hydrogens means more stability.
Steam distillation needs	Compound must be volatile in steam AND immiscible with water (e.g. aniline).

CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 9 : Hydrocarbons (NCERT)

9.1 Classification & General Formulae

Class	General Formula	Example	DoU	Bond
Alkane	C_nH_{2n+2}	Methane CH_4	0	C−C (sp^3)
Alkene	C_nH_{2n}	Ethene C_2H_4	1	C=C (sp^2)
Alkyne	C_nH_{2n-2}	Ethyne C_2H_2	2	C≡C (sp)
Cycloalkane	C_nH_{2n}	Cyclopentane C_5H_{10}	1	C−C ring
Arene (benzene)	C_nH_{2n-6}	Benzene C_6H_6	4	Delocalised

9.2 Combustion (General Equations)

Compound	Combustion Equation
Alkane	$C_nH_{2n+2} + (3n+1)/2\ O_2 \rightarrow n\ CO_2 + (n+1)\ H_2O$
General hydrocarbon	$C_xH_y + (x + y/4)\ O_2 \rightarrow x\ CO_2 + (y/2)\ H_2O$
Methane	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
Ethene	$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$
Ethyne	$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$

9.3 Alkanes — Preparation & Reactions

Reaction	Equation / Condition
Wurtz reaction	2R−X + 2Na → R−R + 2NaX (dry ether)
Hydrogenation (Sabatier)	Alkene/alkyne + H_2 → alkane (Ni/Pt/Pd, Δ)
Decarboxylation	RCOONa + NaOH → R−H + Na_2CO_3 (CaO, Δ)
Kolbe electrolysis	$2RCOONa + 2H_2O \rightarrow R{-}R + 2CO_2 + H_2 + 2NaOH$
Frankland reaction	$2R{-}X + Zn \rightarrow R{-}R + ZnX_2$
Halogenation (free radical)	$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ (UV light); reactivity 3°H > 2°H > 1°H
Free-radical steps	Initiation: $Cl_2 \rightarrow 2Cl\bullet$; Propagation & Termination
Isomerisation / Pyrolysis	n-alkane → branched ($AlCl_3$, HCl) ; cracking → smaller alkane + alkene (high T)

9.4 Alkenes — Reactions & Markovnikov's Rule

Reaction	Equation / Rule
Markovnikov's rule	H^+ adds to C with MORE H atoms (forms more stable carbocation)
Anti-Markovnikov (peroxide / Kharasch)	Only with HBr + peroxide: Br adds to C with MORE H (free-radical)

Reaction	Equation / Rule
Hydrogenation	$CH_2=CH_2 + H_2 \rightarrow CH_3-CH_3$ (Ni/Pt/Pd)
Halogenation	$CH_2=CH_2 + Br_2 \rightarrow BrCH_2-CH_2Br$ (decolourises Br_2 water)
Hydration	$CH_2=CH_2 + H_2O \rightarrow CH_3CH_2OH$ (dil. H_2SO_4)
Ozonolysis (reductive)	Alkene + O_3 → ozonide → (Zn/H_2O) → aldehydes/ketones
Baeyer's test (oxidation)	Cold dilute alkaline $KMnO_4$ → vicinal diol (pink decolourised)
Polymerisation	n $CH_2=CH_2$ → polythene (high P, catalyst)

9.5 Alkynes — Reactions

Reaction	Equation / Condition
Hydrogenation → alkene	$RC\equiv CR + H_2 \rightarrow$ cis-alkene (Lindlar's catalyst: $Pd/BaSO_4$ + quinoline)
Hydrogenation → alkane	$RC\equiv CR + 2H_2 \rightarrow$ alkane (Ni/Pt)
Hydration (Markovnikov)	$HC\equiv CH + H_2O \rightarrow CH_3CHO$ ($HgSO_4$, dil. H_2SO_4)
Acidity of terminal alkyne	$HC\equiv CH + Na \rightarrow HC\equiv C^-Na^+ + \frac{1}{2}H_2$; reacts with $NaNH_2$
Silver acetylide test	Terminal alkyne + ammoniacal $AgNO_3$ → white precipitate
Cuprous acetylide test	Terminal alkyne + ammoniacal Cu_2Cl_2 → red precipitate
Acidity order	$HC\equiv CH > H_2C=CH_2 > CH_3-CH_3$ ($sp > sp^2 > sp^3$; more s-character)
Cyclic polymerisation	3 $HC\equiv CH$ → benzene (red-hot Fe tube, 873 K)

9.6 Aromatic Hydrocarbons — Benzene

Concept	Detail
Aromaticity (Huckel's rule)	Planar, cyclic, fully conjugated, with $(4n + 2)$ π electrons (n = 0,1,2…)
Benzene	6 π electrons (n = 1) → aromatic; all C−C bonds equal (resonance hybrid)
Electrophilic substitution (EAS)	Halogenation, nitration, sulphonation, Friedel–Crafts alkylation & acylation
Nitration	C_6H_6 + conc. HNO_3/conc. H_2SO_4 → nitrobenzene ($E^+ = NO_2^+$)
Halogenation	$C_6H_6 + Cl_2 \rightarrow$ chlorobenzene (anhyd. $FeCl_3/AlCl_3$)
Friedel–Crafts alkylation	$C_6H_6 + RCl \rightarrow C_6H_5R$ (anhyd. $AlCl_3$)
Friedel–Crafts acylation	$C_6H_6 + RCOCl \rightarrow C_6H_5COR$ (anhyd. $AlCl_3$)

9.7 Directing Effects in Benzene (EAS)

Director Type	Groups
Ortho/para directing (activating)	−OH, −OR, $-NH_2$, −NHR, $-NR_2$, −NHCOR, $-CH_3$ (−R) → increase reactivity
Ortho/para directing (deactivating)	−F, −Cl, −Br, −I (halogens: deactivate but direct o/p)
Meta directing (deactivating)	$-NO_2$, −CN, −CHO, −COR, −COOH, $-SO_3H$ → decrease reactivity
Reactivity order	Activated ring > benzene > deactivated ring

9.8 Important Name Reactions

Name Reaction	Equation / Condition
Wurtz	$2R{-}X + 2Na \rightarrow R{-}R + 2NaX$ (dry ether)
Kolbe electrolysis	$2RCOONa + 2H_2O \rightarrow R{-}R + 2CO_2 + H_2 + 2NaOH$
Frankland	$2R{-}X + Zn \rightarrow R{-}R + ZnX_2$
Ozonolysis	Alkene + $O_3 \rightarrow (Zn/H_2O) \rightarrow$ aldehydes/ketones
Friedel–Crafts	Benzene + RCl/RCOCl → alkyl/acyl benzene (anhyd. $AlCl_3$)

★ QUICK MEMORY AIDS & TRICKS

General formulae	Alkane 2n+2, Alkene 2n, Alkyne 2n−2, Arene 2n−6 (each π/ring drops 2 H).
Markovnikov	'Rich get richer' — H adds to the C that already has more H atoms.
Peroxide effect	ONLY HBr + peroxide gives anti-Markovnikov (HCl & HI don't).
Both tests decolourise	Br_2 water AND Baeyer's ($KMnO_4$) decolourise with C=C and C≡C.
Acidity of C−H	$sp > sp^2 > sp^3$ → terminal alkynes are acidic (give metal acetylides).
Huckel's rule	Aromatic if (4n+2) π electrons → 2, 6, 10… (benzene = 6).
o/p vs m directors	Activating groups → o/p ; deactivating (except halogens) → meta.
Lindlar's catalyst	$Pd/BaSO_4$ + quinoline → stops at cis-alkene (poisoned catalyst).