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 = 6.022×10^{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 $1 \le N < 10$ (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 %	$M = 10 \cdot x \cdot d / M_{solute}$ — 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.'