FREE NOTES · NCERT + JEE/NEET LEVEL

Some Basic Concepts of Chemistry

Matter and its classification, SI units, significant figures, laws of chemical combination, Dalton's theory, atomic and molecular mass, the mole concept, percentage composition, empirical formula, stoichiometry, limiting reagent and concentration terms.

01 Matter & measurement
02 Scientific notation & significant figures
03 Laws of chemical combination
04 Atomic & molecular mass
05 The mole concept
06 Percentage composition & empirical formula
07 Stoichiometry & limiting reagent
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01 Matter & measurement

- **Matter:** has mass and occupies space. States: solid, liquid, gas.
- **Pure substances** (elements, compounds; fixed composition) vs **mixtures** (homogeneous: air, sugar solution; heterogeneous: sand + salt).
- **SI base units:** metre (m), kilogram (kg), second (s), kelvin (K), ampere (A), mole (mol), candela (cd).
- Density = mass/volume; T(K) = t(°C) + 273.15; °F = (9/5)°C + 32.

PRECISION VS ACCURACY

Precision = how close repeated readings are to each other. **Accuracy** = how close they are to the true value. Readings 1.95, 1.93, 1.94 for a true value 2.00 are precise but not accurate.

02 Scientific notation & significant figures

Write numbers as $N \times 10^n$ ($1 \le N < 10$): $0.00016 = 1.6 \times 10^{-4}$.

- All non-zero digits count: 285 cm ⇒ 3.
- Zeros between non-zero digits count: 2.005 ⇒ 4.
- Leading zeros don't count: 0.0034 ⇒ 2.
- Trailing zeros count if there is a decimal point: 0.200 ⇒ 3; 100 ⇒ 1 (ambiguous), 100. ⇒ 3.
- Exact numbers (counted items, defined values) have infinite significant figures.

RULES FOR CALCULATION

Add/subtract: answer has as many **decimal places** as the least precise number (12.11 + 18.0 + 1.012 = 31.1). **Multiply/divide:** as many **significant figures** as the number with the fewest (2.5 × 1.25 = 3.1). Rounding: if the dropped digit is 5 and the previous digit is even, leave it (6.35 → 6.4, 6.25 → 6.2).

03 Laws of chemical combination

Law	Statement	By
Conservation of mass	Matter is neither created nor destroyed	Lavoisier
Definite proportions	A compound always has the same elements in the same mass ratio	Proust
Multiple proportions	When two elements form more than one compound, masses of one combining with a fixed mass of the other are in a simple whole-number ratio (H_2O and H_2O_2: O ratio 1:2)	Dalton
Gay Lussac's law of gaseous volumes	Gases combine in simple volume ratios (same T, p): $H_2 + Cl_2 \rightarrow 2HCl$, 1:1:2	Gay Lussac
Avogadro's law	Equal volumes of gases at same T, p contain equal numbers of molecules	Avogadro

Dalton's atomic theory: matter is made of indivisible atoms; atoms of an element are identical; compounds form when atoms combine in fixed ratios; reactions rearrange atoms. (Couldn't explain the law of gaseous volumes.)

04 Atomic & molecular mass

- **1 amu (u)** = 1/12 of the mass of one C-12 atom = 1.66056×10^{-24} g.
- **Average atomic mass** uses isotopic abundance: Cl = (35 × 75.77 + 37 × 24.23)/100 ≈ 35.5 u.
- **Molecular mass** = sum of atomic masses: H_2O = 18 u, $C_{12}H_{22}O_{11}$ = 342 u.
- **Formula mass** for ionic compounds: NaCl = 58.5 u.

05 The mole concept

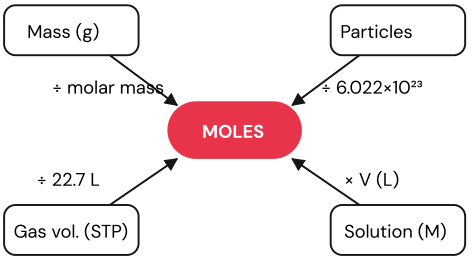
1 mole = the amount containing as many particles as there are atoms in exactly 12 g of C-12 = **6.022×10^{23}** (Avogadro constant N_A).

$$n = \text{mass} / \text{molar mass}$$

$$n = \text{number of particles} / N_A$$

$$n = \text{volume of gas at STP} / 22.7\ \text{L}$$

Molar mass in g/mol = molecular mass in u (H_2O: 18 g/mol). At STP (273.15 K, 1 bar) molar volume is 22.7 L (22.4 L at 1 atm, older convention).



Everything converts through moles

06 Percentage composition & empirical formula

Mass % of an element = (mass of element in 1 mol / molar mass) × 100. In H_2O: H = 2.016/18.02 × 100 = 11.19%.

STEPS FOR EMPIRICAL FORMULA

1. Take % as grams (in 100 g).
2. Divide each by its atomic mass ⇒ moles.
3. Divide by the smallest ⇒ simple ratio (multiply if needed to get whole numbers).
4. Molecular formula = n × empirical formula, n = molar mass / empirical formula mass.

07 Stoichiometry & limiting reagent

A balanced equation gives mole ratios: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ means 1 mol CH_4 needs 2 mol O_2 (16 g with 64 g) and gives 1 mol CO_2 + 2 mol H_2O.

LIMITING REAGENT

The reactant that runs out first and decides how much product forms. Method: divide moles of each reactant by its coefficient; the smallest value is the limiting reagent.

08 Concentration terms

Term	Formula
Mass %	$\text{mass of solute} / \text{mass of solution} \times 100$
Mole fraction	$x_A = n_A/(n_A + n_B)$
Molarity (M)	$\text{moles of solute} / \text{litres of solution}$
Molality (m)	$\text{moles of solute} / \text{kg of solvent}$
Dilution	$M_1V_1 = M_2V_2$

09 JEE Extra + Tricks

JEE LEVEL

- Equivalent concept: equivalents = moles × n-factor; N = M × n-factor. n-factor of acid = replaceable H^+, base = OH^-, oxidant/reductant = electrons gained/lost per molecule.
- Volume strength of H_2O_2 = 11.2 × molarity (at STP); % labelled oleum: % = 100 + (mass of water needed per 100 g).
- POAC (principle of atom conservation): atoms of an element are conserved; saves balancing in stoichiometry.
- Limiting reagent trick: divide moles of each reactant by its coefficient; smallest value limits.
- Molality is temperature independent; molarity changes with temperature (volume changes).
- Molarity from % w/w and density: M = (10 × d × %)/(molar mass).

10 Solved numericals

Q1. How many molecules in 9 g of water?

n = 9/18 = 0.5 mol ⇒ 0.5 × 6.022 × 10^{23} = **3.011 × 10^{23}**.

Q2. A compound has 4.07% H, 24.27% C, 71.65% Cl; molar mass 98.96 g. Find the formulas.

Moles: H 4.04, C 2.02, Cl 2.02 ⇒ ratio 2:1:1 ⇒ empirical CH_2Cl (49.48) ⇒ n = 2 ⇒ **$C_2H_4Cl_2$**.

Q3. 3 mol H_2 react with 2 mol N_2 ($N_2 + 3H_2 \rightarrow 2NH_3$). Limiting reagent and NH_3 formed?

H_2: 3/3 = 1; N_2: 2/1 = 2 ⇒ H_2 limiting ⇒ **2 mol NH_3**.

Q4. 4 g NaOH dissolved to make 250 mL solution. Molarity?

n = 4/40 = 0.1 mol ⇒ M = 0.1/0.25 = **0.4 M**.

11 Mistakes & practice

- Counting leading zeros as significant.
- Using millilitres directly in molarity (convert to litres).
- Picking the reactant with fewer grams as limiting (compare moles ÷ coefficient).
- Mixing up empirical and molecular formula.

1. Significant figures in 0.0025?
2. Mass of 1 molecule of CO_2?
3. Volume of 2 mol O_2 at STP (bar convention)?
4. Which law do CO and CO_2 illustrate?
5. Empirical formula of glucose?

ANSWERS

1) 2 · 2) $44/(6.022 \times 10^{23}) = 7.3 \times 10^{-23}$ g · 3) 45.4 L · 4) Multiple proportions · 5) CH_2O

12 Quick revision: 5 minutes before the exam

MEASUREMENT

- 7 SI base units.
- Leading zeros not significant.
- +/− decimals; ×/÷ sig figs.

MOLE

- $N_A = 6.022 \times 10^{23}$.
- $n = m/M = N/N_A = V/22.7$.
- $1\ u = 1.66 \times 10^{-24}$ g.

LAWS

- Lavoisier, Proust, Dalton.
- Gay Lussac: volume ratios.
- Avogadro: equal V, equal N.

CALCULATIONS

- Empirical: % → moles → ratio.
- Limiting: smallest moles/coeff.
- $M_1V_1 = M_2V_2$.

Have a doubt? Ask the AI Doubt Solver at padhaiadda.netlify.app, or take the mock test on this chapter.