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Some Basic Concepts of Chemistry: Chapter Introduction

1. The hook: how many molecules in one drop?

One drop of water ≈ 0.05 mL ≈ 0.05 g. Molar mass of water = 18 g mol^{-1}.

$n = 0.05 \div 18 = 2.78 \times 10^{-3}$ mol
$N = n \times N_A = 2.78 \times 10^{-3} \times 6.022 \times 10^{23} \approx$ **1.7×10^{21} molecules**

Counting one per second would take about 5×10^{13} years, thousands of times the age of the universe (about 1.4×10^{10} years). Chemists therefore count particles in a bigger unit: the **mole**.

2. Matter

Matter: anything that has mass and occupies space.

State	Particles	Shape / volume
Solid	closely packed, fixed positions	definite shape and volume
Liquid	close but free to move	definite volume, takes the container's shape
Gas	far apart, move freely	no definite shape or volume

Element: one kind of atom (O_2, Fe). **Compound:** two or more elements combined in a fixed ratio (H_2O, CO_2). **Mixture:** substances mixed in any proportion, each keeping its properties (air, salt water).

3. Measurement

Chemistry uses **SI units** (kg, m, s, K, mol …) and reports readings with the right number of **significant figures**, e.g. 25.0 mL (3 significant figures) from a measuring cylinder.

4. Laws that led to atoms

Law of conservation of mass (Lavoisier): in a chemical reaction, mass is neither created nor destroyed. Total mass of reactants = total mass of products.

Other laws of chemical combination: definite proportions, multiple proportions, Gay-Lussac's law of gaseous volumes, Avogadro's law. To explain them, **John Dalton (1808)** proposed that matter is made of tiny indivisible **atoms**; atoms of an element are identical; compounds form when atoms combine in fixed, simple ratios.

5. The mole: a counting unit

1 mole = exactly **$6.022\ 140\ 76 \times 10^{23}$** particles (the Avogadro constant N_A, usually 6.022×10^{23}).

Just as 1 dozen = 12 eggs, 1 mole = 6.022×10^{23} atoms, molecules or ions. The mass of one mole of a substance is its **molar mass** (water: $2 \times 1 + 16 = 18$ g mol^{-1}).

Number of moles: n = mass ÷ molar mass; **number of particles:** $N = n \times N_A$.

6. Stoichiometry

Stoichiometry uses a balanced equation to find how much reactant is needed and how much product forms.

Solved example. $2H_2 + O_2 \rightarrow 2H_2O$. How much O_2 and how much water for 4 mol of H_2?
Mole ratio $H_2 : O_2 : H_2O = 2 : 1 : 2$. So 4 mol H_2 needs $4 \times ½ =$ **2 mol O_2** and gives **4 mol H_2O** ($4 \times 18 = 72$ g).

Key takeaways

- Chemistry starts with matter: states and kinds.
- Measurements use SI units and significant figures.
- Mass is conserved; such laws led Dalton to atoms.
- 1 mole = 6.022×10^{23} particles, a counting unit like a dozen.
- Stoichiometry turns a balanced equation into amounts.

Practice questions (CBSE style)

1. (MCQ) Which is a mixture? (a) water (b) oxygen (c) air (d) carbon dioxide
2. (MCQ) One mole of particles is: (a) 12 (b) 6.022×10^{23} (c) 6.022×10^{22} (d) 1000
3. (Very short) State the law of conservation of mass.
4. (Very short) Find the molar mass of CO_2 (C = 12, O = 16).
5. (Short) Why do chemists count particles in moles?
6. (Numerical) How many molecules are in 9 g of water?
7. (Numerical) In $2H_2 + O_2 \rightarrow 2H_2O$, how many grams of water form from 2 mol of O_2?
8. (Application) Burning paper seems to lose mass. Does this break the law of conservation of mass? Explain.

Answer key

1. (c) air: its gases keep their own properties and can vary in proportion.
2. (b) 6.022×10^{23}.
3. In a chemical reaction mass is neither created nor destroyed; total mass of reactants = total mass of products.
4. $12 + 2 \times 16 = 44$ g mol^{-1}.
5. Atoms and molecules are far too small and too many to count one by one; a mole groups 6.022×10^{23} of them, like a dozen groups 12.
6. n = 9 ÷ 18 = 0.5 mol; $N = 0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$ molecules.
7. $O_2 : H_2O = 1 : 2$, so 2 mol O_2 gives 4 mol $H_2O = 4 \times 18 = 72$ g.
8. No. The carbon dioxide and water vapour formed escape into the air; if they are collected, the total mass of products equals the mass of paper plus oxygen used.

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