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Some Basic Concepts of Chemistry

Class 11 Notes | CBSE Chemistry

Chapter 1

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Some Basic Concepts of Chemistry Class 11 Notes | CBSE Chemistry Chapter 1

Some Basic Concepts of Chemistry is Chapter 1 of CBSE Class 11 Chemistry — and the single most important foundation you will build all year. It teaches you the language of chemistry: how we count atoms we cannot see, how we balance reactions, and how a pinch of substance translates into exact numbers of particles. Master the mole concept here and stoichiometry, solutions, equilibrium, and almost every numerical in JEE, NEET, and your board exam suddenly become solvable.

By the end of these notes you will be able to apply the laws of chemical combination, convert effortlessly between mass, moles and number of particles, find empirical and molecular formulae, solve limiting-reagent problems, and calculate molarity, molality and mole fraction with confidence. This is a high-weightage chapter carrying roughly 6–8 marks in boards, and the gateway to all of physical chemistry.

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Key Concepts

1. Matter and Its Classification

Matter is anything that has mass and occupies space. At the macroscopic level it exists in three physical states — solid, liquid and gas — that can be interconverted by changing temperature and pressure.

Chemically, matter is classified into **mixtures** and **pure substances**. A mixture (e.g., air, sugar solution) contains two or more components in any ratio, while a pure substance (e.g., water, gold) has a fixed composition.

- **Homogeneous mixture:** uniform composition throughout (e.g., salt dissolved in water).
 - **Heterogeneous mixture:** non-uniform composition (e.g., a mixture of sand and iron filings).
 - **Elements:** contain only one kind of atom (e.g., Na, O₂).
 - **Compounds:** two or more elements combined in a fixed ratio by mass (e.g., H₂O, CO₂).
-

2. Properties and Their Measurement (SI Units)

Physical properties (e.g., density, melting point) can be measured without changing the identity of the substance, while chemical properties (e.g., combustibility) involve a chemical change.

Chemistry uses the **SI system** of seven base units. The two you use most often are the **kilogram (kg)** for mass and the **mole (mol)** for amount of substance.

Precision, Accuracy and Significant Figures

- **Accuracy:** how close a measured value is to the true value.
- **Precision:** how close repeated measurements are to one another.
- **Significant figures** are the meaningful digits known with certainty plus one uncertain digit. All non-zero digits are significant; leading zeros are not; trailing zeros after a decimal are significant.

Rules for calculation: In multiplication/division, the answer has as many significant figures as the term with the fewest. In addition/subtraction, the answer has as many decimal places as the term with the fewest.

3. Laws of Chemical Combination

These five laws govern how elements combine to form compounds and laid the experimental basis for Dalton's atomic theory.

The Five Laws

- **Law of Conservation of Mass (Lavoisier):** mass can neither be created nor destroyed in a chemical reaction — total mass of reactants equals total mass of products.
 - **Law of Definite (Constant) Proportions (Proust):** a given compound always contains the same elements in the same fixed proportion by mass (e.g., water is always 1:8 H:O by mass).
 - **Law of Multiple Proportions (Dalton):** when two elements form more than one compound, the masses of one element that combine with a fixed mass of the other are in a simple whole-number ratio (e.g., in CO and CO₂ the oxygen ratio is 1:2).
 - **Gay Lussac's Law of Gaseous Volumes:** when gases react, they do so in volumes that bear a simple whole-number ratio to one another and to the products, at constant T and P.
 - **Avogadro's Law:** equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules.
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4. Dalton's Atomic Theory

In 1808 John Dalton proposed the first modern atomic theory, explaining the laws of chemical combination.

- Matter consists of indivisible particles called **atoms**.
- Atoms of the same element are identical in mass and properties; atoms of different elements differ.
- Atoms combine in simple whole-number ratios to form compounds.

- Atoms are neither created nor destroyed in a chemical reaction.

Key idea: The first two postulates were later modified — atoms are divisible (electrons, protons, neutrons) and isotopes of an element differ in mass.

5. Atomic and Molecular Mass

Atomic mass is the mass of one atom expressed in **atomic mass units (u)**, where $1\text{ u} = 1/12$ the mass of one carbon-12 atom ($1\text{ u} = 1.66 \times 10^{-24}\text{ g}$).

Average atomic mass accounts for the natural abundance of isotopes. For example, chlorine (75% ^{35}Cl , 25% ^{37}Cl) has average atomic mass $= (35 \times 0.75) + (37 \times 0.25) = 35.5\text{ u}$.

- **Molecular mass:** sum of the atomic masses of all atoms in a molecule (e.g., $\text{H}_2\text{O} = 2 \times 1 + 16 = 18\text{ u}$).
 - **Formula mass:** used for ionic compounds that exist as lattices, not molecules (e.g., $\text{NaCl} = 23 + 35.5 = 58.5\text{ u}$).
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6. The Mole Concept and Molar Mass

One **mole** is the amount of substance that contains exactly 6.022×10^{23} elementary entities (atoms, molecules, ions). This number is **Avogadro's number (N_a)**.

The **molar mass** is the mass of one mole of a substance in grams. Numerically it equals the atomic/molecular mass in u (e.g., molar mass of $\text{H}_2\text{O} = 18\text{ g/mol}$).

The Master Conversion Formulae

- **Number of moles, $n = \text{given mass (m)} / \text{molar mass (M)}$**
- **Number of particles $= n \times N_a = (m/M) \times 6.022 \times 10^{23}$**
- For a gas at STP (273.15 K, 1 bar): **volume $= n \times 22.7\text{ L}$** (22.4 L at 1 atm).

Key idea: The mole is the bridge that connects the mass you weigh in the lab to the number of atoms reacting at the particle level.

7. Percentage Composition

The **percentage composition** tells you the mass percent of each element in a compound, useful for checking purity and finding formulae.

Mass % of an element $= (\text{mass of element in 1 mole of compound} / \text{molar mass of compound}) \times 100$

For example, in water ($M = 18$): mass % of H $= (2/18) \times 100 = 11.1\%$, and mass % of O $= (16/18) \times 100 = 88.9\%$.

8. Empirical and Molecular Formula

The **empirical formula** gives the simplest whole-number ratio of atoms in a compound. The **molecular formula** gives the actual number of atoms in one molecule.

Molecular formula = $n \times$ empirical formula, where **$n = \text{molecular mass} / \text{empirical formula mass}$** .

Steps to Find the Empirical Formula

- Convert the mass % of each element to moles (divide by atomic mass).
- Divide all mole values by the smallest one to get the simplest ratio.
- If needed, multiply to obtain whole numbers — these become the subscripts.

For example, a compound that is 40% C, 6.7% H and 53.3% O gives the empirical formula CH_2O ; if its molecular mass is 180, then $n = 180/30 = 6$, so the molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$ (glucose).

9. Stoichiometry and Stoichiometric Calculations

Stoichiometry is the quantitative relationship between reactants and products in a balanced chemical equation. The coefficients give the mole ratio in which species react.

Consider: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$. This reads: 1 mole of CH_4 reacts with 2 moles of O_2 to give 1 mole of CO_2 and 2 moles of H_2O .

Steps for a Stoichiometry Problem

- Write and balance the chemical equation.
 - Convert the given quantity to moles.
 - Use the mole ratio from the equation to find moles of the required species.
 - Convert moles back to mass, volume, or number of particles as asked.
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10. Limiting Reagent

The **limiting reagent** is the reactant that is completely consumed first and therefore decides the maximum amount of product formed. The other reactant(s) are left over in excess.

Method: Find the moles of each reactant, divide each by its stoichiometric coefficient, and the smallest value identifies the limiting reagent.

Key idea: Product yield is always calculated from the limiting reagent, never from the excess reactant — a common board-exam trap.

11. Concentration Terms (Methods of Expressing Concentration)

The concentration of a solution tells you how much solute is dissolved in a given amount of solvent or solution. CBSE focuses on four temperature-independent or temperature-dependent terms.

Term	Definition	Formula
Mass percent (% w/w)	Mass of solute per 100 g of solution	$(\text{mass of solute} / \text{mass of solution}) \times 100$
Molarity (M)	Moles of solute per litre of solution (temperature-dependent)	$M = \text{moles of solute} / \text{volume of solution (L)}$
Molality (m)	Moles of solute per kg of solvent (temperature-independent)	$m = \text{moles of solute} / \text{mass of solvent (kg)}$
Mole fraction (x)	Moles of a component per total moles	$x_A = n_A / (n_A + n_B)$

Note: Molality and mole fraction are independent of temperature because they use mass, not volume; molarity changes with temperature because volume expands on heating. The sum of mole fractions of all components equals 1.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Mole concept, stoichiometry, concentration terms, empirical formula
JEE Main / Advanced	2–3 questions	Limiting reagent, molarity/molality interconversion, % yield
NEET	2–3 questions	Mole concept, significant figures, laws of combination

[TABLE: Question-type split — VSA (1 mark): definitions, laws, units; SA (2–3 marks): mole/mass conversions, concentration calculations, significant figures; LA (5 marks): empirical-to-molecular formula derivations, limiting-reagent numericals.]

Important Definitions

Term	Definition
Mole	Amount of substance containing 6.022×10^{23} elementary entities
Avogadro's number (N_a)	Number of particles in one mole = 6.022×10^{23}
Molar mass	Mass of one mole of a substance in grams
Atomic mass unit (u)	1/12 the mass of one carbon-12 atom
Law of conservation of mass	Mass is neither created nor destroyed in a chemical reaction
Law of definite proportions	A compound always has the same elements in the same mass ratio
Empirical formula	Simplest whole-number ratio of atoms in a compound
Limiting reagent	Reactant fully consumed first, deciding the amount of product
Molarity (M)	Moles of solute per litre of solution
Molality (m)	Moles of solute per kilogram of solvent

Solved Examples

Example 1

Calculate the number of moles and the number of molecules in 36 g of water (H_2O).

Answer: Molar mass of H_2O = 18 g/mol. $n = m/M = 36/18 = 2 \text{ mol}$. Number of molecules = $2 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}$ molecules.

Example 2

How many grams of sodium are present in 0.5 mole of Na? (atomic mass of Na = 23)

Answer: $m = n \times M = 0.5 \times 23 = 11.5 \text{ g}$.

Example 3

A compound contains 4.07% H, 24.27% C and 71.65% Cl. Its molar mass is 98.96 g/mol. Find the molecular formula.

Answer: Moles $\rightarrow$ H: $4.07/1 = 4.07$; C: $24.27/12 = 2.02$; Cl: $71.65/35.5 = 2.02$. Divide by 2.02 $\rightarrow$ H:2, C:1, Cl:1. Empirical formula = CH_2Cl (mass = 49.48). $n = 98.96/49.48 = 2$. **Molecular formula = $C_2H_4Cl_2$.**

Example 4

Calculate the molarity of a solution containing 5 g of NaOH dissolved in 450 mL of solution. (molar mass NaOH = 40)

Answer: Moles of NaOH = $5/40 = 0.125$ mol. Volume = 0.450 L. $M = 0.125/0.450 = \mathbf{0.278 \text{ mol/L}}$.

Example 5

In the reaction $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, if 4 mol of N_2 react with 9 mol of H_2 , identify the limiting reagent and the moles of NH_3 formed.

Answer: N_2 : $4/1 = 4$; H_2 : $9/3 = 3$. H_2 has the smaller value, so **H_2 is the limiting reagent**. NH_3 formed = $9 \times (2/3) = \mathbf{6 \text{ mol}}$.

Example 6

Calculate the mass percent of carbon in carbon dioxide (CO_2).

Answer: Molar mass of $\text{CO}_2 = 12 + (2 \times 16) = 44$. Mass % of C = $(12/44) \times 100 = \mathbf{27.27\%}$.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. Define one mole of a substance.
2. State the law of conservation of mass.
3. How many significant figures are present in 0.00250?
4. What is the SI unit of amount of substance?
5. Why is molality preferred over molarity in temperature-dependent experiments?

2–3-Mark Questions (SA)

1. State the law of multiple proportions and illustrate it with the oxides of carbon (CO and CO_2).
2. Calculate the number of atoms in 0.1 mole of H_2SO_4 .
3. Distinguish between empirical formula and molecular formula with one example each.
4. Define molarity and mole fraction, and state which is temperature-independent and why.

5-Mark Questions (LA)

1. State and explain the five laws of chemical combination with one example for each.
2. A compound has 40% C, 6.7% H and 53.3% O with molar mass 180. Derive its empirical and molecular formula.
3. Explain the concept of limiting reagent. For the reaction $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, if 10 g of H_2 reacts with 64 g of O_2 , find the limiting reagent and the mass of water produced.

Quick Revision Points

- Matter → mixtures (homogeneous/heterogeneous) and pure substances (elements/compounds)
- Five laws of combination: conservation of mass, definite proportions, multiple proportions, Gay Lussac, Avogadro
- 1 mole = 6.022×10^{23} particles = Avogadro's number (N_a)
- $n = m/M$; number of particles = $(m/M) \times N_a$; gas at STP occupies 22.7 L/mol
- Molar mass (g/mol) is numerically equal to molecular mass (u)
- Mass % of element = $(\text{mass of element} / \text{molar mass}) \times 100$
- Empirical = simplest ratio; molecular = $n \times \text{empirical}$, $n = \text{molecular mass} / \text{empirical mass}$
- Limiting reagent decides product yield — divide moles by coefficients, pick the smallest
- Molarity = mol/L (temperature-dependent); molality = mol/kg solvent (temperature-independent)
- Mole fraction $x_A = n_A / (n_A + n_B)$; sum of all mole fractions = 1
- Significant figures: multiplication/division → fewest sig figs; addition/subtraction → fewest decimal places

Next Chapter: Chapter 2 — Structure of Atom



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