CHAPTER 1

Some Basic Concepts of Chemistry

CBSE Class 11 · Chemistry · Chapter 1

CBSE · Chemistry · Class 11

WHAT THIS CHAPTER DOES

A Convert fluently between mass, moles, and number of particles using Avogadro's number.

B State and apply the five laws of chemical combination.

Boards prep that builds confidence, not anxiety.

TODAY'S MISSION

Today's mission

1. Convert fluently between mass, moles, and number of particles using Avogadro's number.
2. State and apply the five laws of chemical combination.
3. Calculate molarity, molality, and mole fraction of any solution.
4. Identify the limiting reagent and compute the exact mass of product formed.

WHY THIS MATTERS

Why this chapter matters

1. 6-8 marks every Class 11 paper — and the conceptual bedrock for ALL of Class 11 and 12 Physical Chemistry.
2. The mole concept is the single most-used tool in chemistry; every later numerical depends on it.
3. Real-world bridge — dosage calculations, fertiliser yields, solution-making in any lab all run on these exact formulae.

TOPIC

A

Nature of matter and the laws of chemical combination

TOPIC

Matter and its classification

STATES OF MATTER

Matter is anything that has mass and occupies space, and it exists in three physical states. In a SOLID the particles are tightly packed in a fixed arrangement, so a solid has a definite shape and a definite volume. In a LIQUID the particles are close but mobile, giving a definite volume but the shape of the container. In

ELEMENT / COMPOUND / MIXTURE

A pure substance has a fixed composition. An ELEMENT is the simplest pure substance, made of only one kind of atom (e.g. Na, O2, gold). A COMPOUND is a pure substance formed when two or more elements combine in a FIXED mass ratio, with completely new properties (water

PHYSICAL VS CHEMICAL PROPERTIES

A PHYSICAL property can be measured or observed without changing the substance's chemical identity — colour, density, melting point, boiling point, solubility and physical state are all physical properties, and a physical change (ice melting, sugar dissolving) is

MEASUREMENT AND UNITS

Chemistry is a quantitative science, so every measurement carries a number AND a unit, expressed in the SI system: mass in kilograms (g in practice), length in metres, amount of substance in MOLES, and temperature in kelvin (K = degrees C + 273.15). Two ideas govern reporting:

THEOREM · LOAD-BEARING RESULT

The Five Laws of Chemical Combination

" 1. Conservation of mass (Lavoisier): in a chemical reaction, total mass of reactants = total mass of products. 2. Constant proportions (Proust): a given compound always contains the same elements in the same fixed mass ratio. 3. Multiple proportions (Dalton): when two elements form more than one compound, the masses of one element combining with a fixed mass of the other are in a simple

STATEMENT

4. Gay-Lussac's law of gaseous volumes: gases react and form products in volume ratios that are simple whole numbers, at the same temperature and pressure. 5. Avogadro's law: equal

WHY THIS MATTERS

- These five laws are the experimental evidence that matter is made of discrete atoms combining in fixed ratios — they are precisely what Dalton's atomic theory was invented to explain.

WATCH OUT FOR

NOTE Do NOT confuse Gay-Lussac's law (volume RATIOS of reacting gases) with Avogadro's law (equal volumes contain equal NUMBERS of molecules). Avogadro's law is the deeper reason Gay-Lussac's ratios come out whole.

TOPIC

Dalton's atomic theory and atomic masses

DALTON'S POSTULATES

John Dalton (1808) proposed that all matter is made of tiny indivisible particles called atoms; that atoms of a given element are identical in mass and properties while atoms of different elements differ; that compounds form when atoms of different elements combine in fixed, simple whole-

RELATIVE ATOMIC MASS AND THE UNIT U

Atoms are far too light to weigh individually, so chemists use RELATIVE atomic mass, comparing each atom against a standard. The modern standard is the carbon-12 isotope, defined as exactly 12 atomic mass units (amu, symbol u). One unified atomic mass unit u =

AVERAGE ATOMIC MASS

Most elements occur naturally as a mixture of ISOTOPES — atoms of the same element with different numbers of neutrons and hence different masses. The atomic mass shown in the periodic table is the AVERAGE atomic mass, the abundance-weighted mean over all natural isotopes: average = sum

MOLECULAR AND FORMULA MASS

The MOLECULAR MASS of a substance is the sum of the atomic masses of all atoms in one molecule: for water H_2O it is 2(1.008) + 16.00 = 18.02 u, and for methane CH_4 it is 12.01 + 4(1.008) = 16.04 u. For IONIC compounds, which have no discrete molecules, we instead

TOPIC

B

The mole concept and molar mass

THEOREM · LOAD-BEARING RESULT

The Mole and Avogadro's Number

> A MOLE is the amount of a substance that contains exactly 6.022 x 10^23 elementary entities (atoms, molecules, ions, or formula units). This number, 6.022 x 10^23 per mole, is Avogadro's number, N_A. One mole is simply a chemist's counting unit, just as 'a dozen' means 12.

STATEMENT

The MOLAR MASS of a substance is the mass of one mole, numerically equal to its atomic / molecular / formula mass but expressed in grams per mole (g/mol). Thus 1 mol of carbon-12 weighs exactly

WHY THIS MATTERS

- Atoms react in whole-number ratios but are far too small to count or weigh singly
- The mole bridges the invisible atomic scale to the laboratory gram scale, letting us 'count by weighing'.

WATCH OUT FOR

NOTE A mole is a COUNT, not a mass. 1 mol H2 (2 g) and 1 mol O2 (32 g) have very different masses but the SAME number of molecules. Never assume a mole has a fixed weight.

WORKED EXAMPLE

Moles, molecules and atoms in 36 g of water

1. Molar mass of H2O = 2(1.008) + 16.00 = 18.02 g/mol (use 18 g/mol).
2. Number of moles n = mass / molar mass = 36 g / 18 g/mol = 2 mol.
3. Number of molecules = n x N_A = 2 x 6.022 x 10^23 = 1.2044 x 10^24 molecules.
4. Each H2O has 3 atoms, so number of atoms = 1.2044 x 10^24 x 3 = 3.6132 x 10^24 atoms.

WORKED EXAMPLE

Percentage composition of water by mass

1. Molar mass of H2O = 2(1.008) + 16.00 = 18.016 g/mol.
2. Mass of hydrogen per mole = 2 x 1.008 = 2.016 g; mass of oxygen = 16.00 g.
3. % H = (2.016 / 18.016) x 100 = 11.19%.
4. % O = (16.00 / 18.016) x 100 = 88.81%. Check: 11.19 + 88.81 = 100.00% ✓

TRY IT · SOLVE BEFORE YOU PEEK

How many molecules are present in 8 g of oxygen gas (O_2)? (Take molar mass O_2 = 32 g/mol.)

SOLUTION

ANSWER n = 8/32 = 0.25 mol. Molecules = $0.25 \times 6.022 \times 10^{23} = 1.5055 \times 10^{23}$ molecules of O_2.

TOPIC

C

Empirical and molecular formula

WORKED EXAMPLE

Empirical and molecular formula from % composition

1. A compound is 40.0% C, 6.7% H, 53.3% O by mass; molar mass = 180 g/mol. Take a 100 g sample so % become grams.
2. Moles: C = 40.0/12 = 3.33; H = 6.7/1 = 6.7; O = 53.3/16 = 3.33.
3. Divide by the smallest (3.33): C = 1, H = 2.01 ~ 2, O = 1. Empirical formula = CH_2O.
4. Empirical mass = 12 + 2(1) + 16 = 30. n = molar mass / empirical mass = 180/30 = 6. Molecular formula = (CH_2O) x 6 = $C_6H_{12}O_6$ (glucose).

TOPIC

Empirical vs molecular formula

EMPIRICAL FORMULA

The empirical formula gives the SIMPLEST whole-number RATIO of atoms of each element in a compound, obtained from experimental percentage composition. For glucose the empirical formula is CH2O, meaning carbon, hydrogen and oxygen are present in the ratio 1:2:1. The empirical formula

MOLECULAR FORMULA

The molecular formula gives the ACTUAL number of atoms of each element in one molecule of the compound. Glucose has the molecular formula C6H12O6 — six carbon, twelve hydrogen, six oxygen atoms per molecule. It is always a whole-number multiple of the empirical formula:

HOW THEY RELATE

Empirical and molecular formulae describe the SAME substance at different levels of detail — the molecular formula is the empirical formula scaled up by the integer n. When n = 1 the two formulae are identical, as for water (H2O) or carbon dioxide (CO2). To go from empirical to molecular you must know

WORKED RELATIONSHIP

Consider a hydrocarbon that is 92.3% C and 7.7% H with molar mass 78 g/mol. Moles: C = 92.3/12 = 7.69, H = 7.7/1 = 7.7; dividing by the smaller gives C:H = 1:1, so the empirical formula is CH with empirical mass 13. Then n = 78/13 = 6, so the molecular formula is C6H6 — benzene. The same percentage

TOPIC

D

Stoichiometry and the limiting reagent

WORKED EXAMPLE

Limiting reagent in $N_2 + 3\,H_2 \rightarrow 2\,NH_3$

1. Balanced equation: $N_2 + 3\,H_2 \rightarrow 2\,NH_3$. Given 28 g N_2 and 10 g H_2.
2. Moles: $n(N_2) = 28/28 = 1$ mol; $n(H_2) = 10/2 = 5$ mol.
3. Divide by coefficients: $N_2 \rightarrow 1/1 = 1.00$; $H_2 \rightarrow 5/3 = 1.67$. Smaller quotient (1.00) means N_2 is the limiting reagent.
4. Product from limiting reagent: 1 mol $N_2 \rightarrow$ 2 mol NH_3 = 2 x 17 = 34 g NH_3. H_2 used = 3 mol (6 g), so H_2 left over = 10 - 6 = 4 g excess.

TOPIC

Four ways to express concentration

MASS PERCENT

Mass percent expresses the mass of solute as a percentage of the total mass of solution: mass % of A = (mass of A / total mass of solution) x 100. For example, a 10% by mass glucose solution contains 10 g of glucose in every 100 g of solution (i.e. 10 g glucose + 90 g water). Mass percent is temperature-independent

MOLE FRACTION

Mole fraction is the ratio of the moles of one component to the TOTAL moles of all components: $x_A = n_A / (n_A + n_B + \ldots)$, and the mole fractions of all components always sum to 1. For a solution of 2 mol ethanol in 8 mol water, x(ethanol) = 2/10 = 0.2 and x(water) = 8/10 = 0.8, summing to

MOLARITY (M)

Molarity is the most-used laboratory concentration unit: M = moles of solute / volume of SOLUTION in litres, with units mol/L. To prepare 0.4 M NaOH you dissolve 0.4 mol (16 g) of NaOH and make the volume up to exactly 1 litre of solution. Because molarity is defined per litre of SOLUTION and

MOLALITY (M)

Molality is moles of solute per kilogram of SOLVENT: m = moles of solute / mass of solvent in kg, with units mol/kg. To make a 1 molal solution you dissolve 1 mol of solute in exactly 1 kg of solvent. Because molality is defined using MASS of solvent, and mass does not change with temperature

WORKED EXAMPLE

Molarity of a NaOH solution

1. Dissolve 4.0 g of NaOH in water to make 250 mL of solution. Molar mass NaOH = 23 + 16 + 1 = 40 g/mol.
2. Moles of NaOH = 4.0 / 40 = 0.10 mol.
3. Volume of solution = 250 mL = 0.250 L.
4. Molarity = moles / litres = 0.10 / 0.250 = 0.40 mol/L = 0.40 M.

TRY IT · SOLVE BEFORE YOU PEEK

0.5 mol of glucose is dissolved in 500 g of water. What is the molality of the solution?

SOLUTION

ANSWER Molality = moles of solute / kg of solvent = 0.5 mol / 0.500 kg = 1.0 mol/kg = 1.0 m. (Note: molality uses mass of SOLVENT, and is temperature-independent.)

TOPIC

Molarity vs molality temperature dependence

TRAP → TRUTH

× **MISTAKE** Molarity and molality are the same thing in different units, and both stay constant when the solution is heated.

✓ **CORRECT** Molarity (M = mol solute / litre of SOLUTION) depends on VOLUME, which expands on heating, so molarity DECREASES as temperature rises. Molality (m = mol solute / kg of SOLVENT) depends only on MASS, which is temperature-independent, so molality stays constant. This is exactly why molality is preferred in thermodynamic and colligative-property work.

TOPIC

Limiting reagent identification

TRAP → TRUTH

× **MISTAKE** The reactant present in the smaller MASS (or fewer grams) is always the limiting reagent.

✓ **CORRECT** Limiting reagent is decided by MOLES compared against the balanced equation's coefficient ratio, never by raw mass. Convert each reactant to moles, divide each by its coefficient, and the SMALLEST quotient is the limiting reagent. A reactant can be heavier in grams yet still be limiting if its molar amount per coefficient is lower.

TOPIC

Average atomic mass vs mass number

TRAP → TRUTH

× **MISTAKE** The average atomic mass of chlorine is 35 because its mass number is 35.

✓ **CORRECT** Mass number (an integer, e.g. 35 for the Cl-35 nuclide) counts protons + neutrons in ONE isotope. Average atomic mass (35.45 u for Cl) is the abundance-WEIGHTED mean over all naturally occurring isotopes, so it is generally NOT an integer. $35.45 = 35 \times 0.7577 + 37 \times 0.2423$.

TOPIC

Mole as a number vs mole as a mass

TRAP → TRUTH

× **MISTAKE** One mole always weighs the same, so a mole of hydrogen and a mole of oxygen have equal mass.

✓ **CORRECT** One mole is a COUNT: exactly 6.022 x 10^23 particles, like 'a dozen' is 12. The MASS of a mole equals the molar mass, which differs per substance: 1 mol H2 = 2 g but 1 mol O2 = 32 g, although BOTH contain 6.022 x 10^23 molecules.

TOPIC

Empirical vs molecular formula

TRAP → TRUTH

× **MISTAKE** The empirical formula is just the molecular formula written with smaller numbers; they describe different molecules.

✓ **CORRECT** Empirical formula gives the SIMPLEST whole-number RATIO of atoms; molecular formula gives the ACTUAL number of atoms in one molecule. They describe the SAME substance. Molecular = (empirical) x n, where n = molar mass / empirical formula mass. For glucose: empirical CH2O, molecular C6H12O6, n = 180/30 = 6.

TOPIC

Significant figures in chemical calculations

TRAP → TRUTH

× **MISTAKE** Always write every digit your calculator shows; more decimals means a more correct answer.

✓ **CORRECT** The result of a multiplication/division carries the SAME number of significant figures as the least-precise data value. 4.0 g / 18.015 g/mol must be reported as 0.22 mol (2 sig figs), not 0.2220. Reporting spurious extra digits is a marked error in school papers; carry extra digits only in intermediate steps, then round at the end.

TOPIC

Gay-Lussac vs Avogadro's law

TRAP → TRUTH

× **MISTAKE** Gay-Lussac's law of gaseous volumes and Avogadro's law say the same thing.

✓ **CORRECT** Gay-Lussac's law (1808) is about gases REACTING: the volumes of gaseous reactants and products bear a simple WHOLE-NUMBER ratio (at same T, P). Avogadro's law (1811) is the deeper EXPLANATION: equal volumes of all gases at the same T and P contain equal NUMBERS of molecules. Avogadro's law explains WHY Gay-Lussac's volume ratios come out as small whole numbers.

TOPPER TEMPLATE · MARK-BY-MARK

3-mark numerical: 'Calculate the number of moles / molecules / atoms in a given mass of a

1. **WRITE THE MOLAR MASS WITH WORKING** — 1 m

   State the formula's molar mass with the addition shown: 'M(CO2) = 12 + 2 x 16 = 44 g/mol'. Showing the arithmetic earns the mark even if a later step slips.

2. **APPLY MOLES = MASS / MOLAR MASS** — 1 m

   Write the formula explicitly, substitute, and compute: 'n = mass / M = 8.8 g / 44 g/mol = 0.2 mol'. Always carry the unit through.

3. **SCALE TO THE ASKED QUANTITY USING AVOGADRO'S NUMBER** — 1 m

   If molecules/atoms are asked, multiply by 6.022 x 10^23: 'Number of molecules = 0.2 x 6.022 x 10^23 = 1.2044 x 10^23'. For atoms, multiply further by atoms-per-molecule. State the final answer with correct significant figures.

TOPPER TEMPLATE · MARK-BY-MARK

5-mark numerical: 'Identify the limiting reagent and calculate the mass of product formed.'

1 — WRITE AND CHECK THE BALANCED EQUATION (1 m)

Write the balanced equation FIRST, e.g. 'N2 + 3 H2 -> 2 NH3'. The mole ratios you read here are the spine of the whole answer; an unbalanced equation forfeits later marks.

2 — CONVERT EVERY REACTANT MASS TO MOLES, THEN DIVIDE BY COEFFICIENT (2 m)

Compute moles of each reactant: 'n(N2) = 28/28 = 1 mol; n(H2) = 10/2 = 5 mol.' Divide each by its coefficient: N2 -> 1/1 = 1; H2 -> 5/3 = 1.67. The SMALLER quotient (1, for N2) marks the limiting reagent.

3 — STATE THE LIMITING REAGENT EXPLICITLY WITH REASON (1 m)

'N2 is the limiting reagent because its coefficient-scaled mole quantity (1.00) is less than that of H2 (1.67); N2 runs out first and decides how much product forms.'

4 — COMPUTE PRODUCT FROM THE LIMITING REAGENT ONLY (1 m)

Use the limiting reagent ratio: 1 mol N2 -> 2 mol NH3 = 2 x 17 = 34 g NH3. Never compute product from the excess reagent. State the leftover excess if asked.

TOPPER TEMPLATE · MARK-BY-MARK

3-mark numerical: 'A compound contains X% A, Y% B, Z% C by mass. Find the empirical formula.'

1	**CONVERT EACH PERCENTAGE TO MOLES (DIVIDE BY ATOMIC MASS)** 1 m	Assume 100 g sample so percentages become grams. Divide each by its atomic mass: 'C: 40/12 = 3.33; H: 6.7/1 = 6.7; O: 53.3/16 = 3.33.'
2	**DIVIDE ALL MOLE VALUES BY THE SMALLEST** 1 m	Smallest is 3.33: 'C = 3.33/3.33 = 1; H = 6.7/3.33 = 2; O = 3.33/3.33 = 1.' Round to nearest whole numbers (use a x2/x3 multiplier only if a value is near .5 or .33).
3	**WRITE THE EMPIRICAL FORMULA (AND MOLECULAR FORMULA IF M IS GIVEN)** 1 m	Empirical formula = CH2O. If molar mass is given, n = M / empirical-mass; e.g. for M = 180, n = 180/30 = 6, so molecular formula = C6H12O6.

PYQ PATTERNS

Top PYQ patterns to drill

#1	Calculate the number of moles / molecules / atoms in a given mass of a substance. (2-3 marks)	Almost every paper (2019, 2020, 2021, 2022, 2023, 2024)
#2	Calculate the molarity / molality of a solution given mass of solute and volume/mass of solution. (3 marks)	2020 SQP, 2021, 2022, 2023 annual
#3	Identify the limiting reagent and calculate the mass of product formed. (3-5 marks)	2019, 2021, 2022, 2024
#4	Determine the empirical formula (and then molecular formula) from percentage composition. (3 marks)	2020, 2022, 2023
#5	State a named law of chemical combination and illustrate / verify it with given data. (2-3 marks)	2019, 2021, 2023

RECAP · MEMORISE THESE

5-line revision

1 Mole bridge — n = mass/M = particles/ (6.022 x 10^23). The mole is a COUNT, not a fixed mass; molar mass (g/mol) is numerically equal to atomic/molecular mass.

2 Formulae — Empirical = simplest atom ratio; molecular = (empirical) x n where n = M / empirical mass. % composition fixes only the ratio.

3 Stoichiometry + concentration — Limiting reagent decided by coefficient-scaled MOLES, not grams. Molarity is per litre of solution (temp-dependent); molality is per kg of solvent (temp-independent).

WHAT'S NEXT

What's next

→

- → Chapter 2 — Structure of Atom (builds on atomic mass and isotopes introduced here).
- → Sit the 15-MCQ Quick Drill (companion PDF) — under 25 minutes, target >= 12/15.
- → Then the full annual-pattern Paper — 30 marks, 60 minutes, with full worked solutions.

You now own the counting toolkit of chemistry.

Prove your mole fluency. Take the drill, sit the paper, master the chapter.

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