ANSWER KEY & MARKING SCHEME · CBSE CLASS 11

Some Basic Concepts of Chemistry

Chemistry · Chapter 1 · Use this with the Board Paper · Companion to Quick Drill

HOW TO USE

Attempt the Board Paper first (closed-book, full time). Then come here. For 2-mark+ questions, compare your answer to the model. For 3-4 mark questions, also consult the **Topper Templates** below — these show the exact step-by-step structure that scores full marks per CBSE marking-scheme conventions.

MODEL ANSWERS · BOARD PAPER

Section A — Very Short Answer (1 mark each, 4 Qs)

Q1. Define one mole. **[1 mark]**

Ans: One mole is the amount of substance containing exactly 6.022 x 10^23 elementary entities (atoms, molecules, ions or formula units).

Q2. State the law of constant proportions. **[1 mark]**

Ans: A given chemical compound always contains the same elements combined together in the same fixed proportion by mass, regardless of its source or method of preparation.

Q3. Why is the atomic mass of chlorine 35.45 u and not a whole number? **[1 mark]**

Ans: Because it is the abundance-weighted AVERAGE of its two natural isotopes (75.77% Cl-35 and 24.23% Cl-37): 35 x 0.7577 + 37 x 0.2423 = 35.45 u.

Q4. Which concentration term, molarity or molality, is independent of temperature, and why? **[1 mark]**

Ans: Molality, because it is defined per kilogram of solvent (mass), and mass does not change with temperature; molarity uses volume, which expands on heating.

Section B — Short Answer I (2 marks each, 3 Qs)

Q5. Calculate the number of moles and the number of molecules in 22 g of carbon dioxide (CO2). **[2 marks]**

Ans: Molar mass CO2 = 12 + 2(16) = 44 g/mol. Moles = 22/44 = 0.5 mol. Number of molecules = 0.5 x 6.022 x 10^23 = 3.011 x 10^23 molecules.

Q6. Distinguish between molarity and molality, giving the defining formula of each. **[2 marks]**

Ans: Molarity (M) = moles of solute / litre of SOLUTION (unit mol/L), and is temperature-dependent. Molality (m) = moles of solute / kilogram of SOLVENT (unit mol/kg), and is temperature-independent.

Q7. State Avogadro's law and explain how it accounts for Gay-Lussac's law of gaseous volumes. **[2 marks]**

Ans: Avogadro's law: equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules. Since volume is proportional to the number of molecules, the simple whole-number VOLUME ratios that Gay-Lussac observed for reacting gases are a direct consequence of the simple whole-number MOLECULE (mole) ratios in the balanced equation.

Section C — Short Answer II (3 marks each, 3 Qs)

Q8. A 4.0 g sample of NaOH is dissolved in water to make 250 mL of solution. Calculate the molarity of the solution. (Molar mass NaOH = 40 g/mol.) **[3 marks]**

Ans: Moles of NaOH = mass / molar mass = 4.0 / 40 = 0.10 mol. Volume of solution = 250 mL = 0.250 L. Molarity = moles / litres = 0.10 / 0.250 = 0.40 mol/L = 0.40 M.

Q9. A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen by mass. Determine its empirical formula. (Atomic masses: C=12, H=1, O=16.) **[3 marks]**

Ans: Take 100 g sample. Moles: C = 40.0/12 = 3.33; H = 6.7/1 = 6.7; O = 53.3/16 = 3.33. Divide by smallest (3.33): C = 1, H = 2.01 ~ 2, O = 1. Empirical formula = CH2O.

Q10. State the law of conservation of mass and the law of multiple proportions, giving one illustration each. **[3 marks]**

> **Ans:** Law of conservation of mass (Lavoisier): in a chemical reaction, total mass of reactants equals total mass of products — e.g. when 12 g C burns with 32 g O_2, exactly 44 g CO_2 forms (12 + 32 = 44). 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 CO_2, the masses of oxygen per 12 g of carbon are 16 g and 32 g, a 1:2 ratio.

Section D — Case Study / Long Answer (Qs as marked, 2 Qs)

Q11. Ammonia is manufactured industrially by the Haber process: $N_2 + 3 H_2 \rightarrow 2 NH_3$. In a trial run, 28 g of nitrogen gas is mixed with 10 g of hydrogen gas. (a) Calculate the moles of each reactant. (b) Identify the limiting reagent, with reasoning. (c) Calculate the maximum mass of ammonia that can be formed. (d) Calculate the mass of the excess reagent left unreacted. (Atomic masses: N=14, H=1.) **[6 marks]**

> **Ans:** (a) Moles of N_2 = 28 / 28 = 1 mol; moles of H_2 = 10 / 2 = 5 mol. (b) Divide each by its coefficient: N_2 -> 1/1 = 1.00; H_2 -> 5/3 = 1.67. The smaller quotient is for N_2, so N_2 is the LIMITING reagent (it runs out first and decides the product). (c) From the equation, 1 mol N_2 gives 2 mol NH_3 = 2 x 17 = 34 g of ammonia. (d) H_2 actually consumed = 3 mol (for 1 mol N_2) = 6 g; H_2 supplied = 10 g; excess H_2 left = 10 - 6 = 4 g.

Q12. Glucose is the simplest sugar used by the body for energy. A sample is found to contain 40.0% C, 6.7% H and 53.3% O by mass, and its molar mass is determined to be 180 g/mol. (a) Determine the empirical formula. (b) Calculate the empirical formula mass. (c) Determine the molecular formula. (d) State the difference between an empirical and a molecular formula. (Atomic masses: C=12, H=1, O=16.) **[5 marks]**

> **Ans:** (a) Moles in 100 g: C = 40.0/12 = 3.33, H = 6.7/1 = 6.7, O = 53.3/16 = 3.33; dividing by 3.33 gives C:H:O = 1:2:1, so empirical formula = CH_2O. (b) Empirical formula mass = 12 + 2(1) + 16 = 30 g/mol. (c) n = molar mass / empirical mass = 180/30 = 6; molecular formula = (CH_2O) x 6 = $C_6H_{12}O_6$. (d) The empirical formula gives only the simplest whole-number RATIO of atoms, while the molecular formula gives the ACTUAL number of atoms of each element in one molecule; molecular = (empirical) x n.

★ TOPPER ANSWER TEMPLATES

3 TEMPLATES · MEMORISE THE FORMAT

★ TOPPER TEMPLATE — 3-mark numerical: 'Calculate the number of moles / molecules / atoms in a given mass of a substance.'

Almost every annual paper

Step 1 [1 mark]	**Write the molar mass with working**	State the formula's molar mass with the addition shown: 'M(CO_2) = 12 + 2 x 16 = 44 g/mol'. Showing the arithmetic earns the mark even if a later step slips.
Step 2 [1 mark]	**Apply moles = mass / molar mass**	Write the formula explicitly, substitute, and compute: 'n = mass / M = 8.8 g / 44 g/mol = 0.2 mol'. Always carry the unit through.
Step 3 [1 mark]	**Scale to the asked quantity using Avogadro's number**	If molecules/atoms are asked, multiply by 6.022×10^{23}: 'Number of molecules = 0.2 x 6.022×10^{23} = 1.2044×10^{23}'. For atoms, multiply further by atoms-per-molecule. State the final answer with correct significant figures.

COMMON LOSS OF MARKS:

- Using the wrong molar mass (e.g. 16 instead of 32 for O_2) destroys the whole numerical.
- Forgetting to multiply by atoms-per-molecule when ATOMS (not molecules) are asked.
- Quoting an answer to 6 calculator digits when the data justifies only 2-3 significant figures.

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

2019, 2021, 2022, 2024

Step 1 [1 mark]	**Write and check the balanced equation**	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.
Step 2 [2 marks]	**Convert every reactant mass to moles, then divide by coefficient**	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.
Step 3 [1 mark]	**State the limiting reagent explicitly with reason**	'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.'
Step 4 [1 mark]	**Compute product from the LIMITING reagent only**	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.

COMMON LOSS OF MARKS:

- Deciding the limiting reagent by comparing grams instead of coefficient-scaled moles.
- Calculating product mass from the EXCESS reagent (the classic trap).
- Forgetting to multiply moles of product by its molar mass to report a MASS.

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

2020, 2022, 2023

Step 1 [1 mark]	**Convert each percentage to moles (divide by atomic mass)**	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.'
Step 2 [1 mark]	**Divide all mole values by the smallest**	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).
Step 3 [1 mark]	**Write the empirical formula (and molecular formula if M is given)**	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.

COMMON LOSS OF MARKS:

- Dividing by the wrong smallest mole value, or not dividing by the smallest at all.
- Rounding 2.5 or 1.33 down instead of multiplying the whole set by 2 or 3 to clear the fraction.
- Stopping at the empirical formula when the molar mass is given and the molecular formula was asked.

MARKING SCHEME — GENERAL NOTES

- In numerical questions, marks are split between method (correct formula + substitution) and the final answer — a correct method with an arithmetic slip still earns method marks.
- Limiting-reagent questions: identifying the limiting reagent by comparing grams instead of coefficient-scaled moles caps the answer at half marks.
- Molar mass must be shown as an explicit sum (e.g. CO2 = 12 + 32 = 44) to earn the molar-mass mark.
- Final answers reported to a grossly excessive number of significant figures (full calculator readout) lose 0.5 mark for presentation.
- For empirical-formula questions, the molecular formula is required ONLY when a molar mass is supplied; stopping at the empirical formula then caps the marks.