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CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 1: Some Basic Concepts of Chemistry – Free PDF Download

Question: Calculate the molar mass of the following: (i) H2O (ii) CO2 (iii) CH4

Contents

Chapter 1, “Some Basic Concepts of Chemistry,” lays the numerical and conceptual foundation of Class 11 Chemistry — atoms, molecules, moles, stoichiometry, and the mole concept — through which every later chapter’s calculations are built. Below is the complete, verified set of in-text and end-of-chapter exercise questions (1.1 to 1.36) from the current NCERT textbook (Reprint 2026-27), each solved with full working.

NCERT Exercise Solutions

Question 1.1

Question: Calculate the molar mass of the following: (i) H_2O (ii) CO_2 (iii) CH_4

Solution: Using atomic masses H = 1.008 u, C = 12.011 u, O = 16.00 u:
(i) Molar mass of H_2O = 2(1.008) + 16.00 = 2.016 + 16.00 = 18.016 ≈ **18.02 g mol^{-1}**
(ii) Molar mass of CO_2 = 12.011 + 2(16.00) = 12.011 + 32.00 = 44.011 ≈ **44.01 g mol^{-1}**
(iii) Molar mass of CH_4 = 12.011 + 4(1.008) = 12.011 + 4.032 = 16.043 ≈ **16.04 g mol^{-1}**

Question 1.2

Question: Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

Solution: Molar mass of Na_2SO_4 = 2(22.99) + 32.06 + 4(16.00) = 45.98 + 32.06 + 64.00 = 142.04 g mol^{-1}
Mass % of Na = (45.98 / 142.04) × 100 = **32.37%**
Mass % of S = (32.06 / 142.04) × 100 = **22.57%**
Mass % of O = (64.00 / 142.04) × 100 = **45.06%**
(Check: 32.37 + 22.57 + 45.06 ≈ 100%)

Question 1.3

Question: Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.

Solution: Moles of Fe = 69.9 / 55.85 = 1.251
Moles of O = 30.1 / 16.00 = 1.881
Dividing by the smaller value (1.251): Fe = 1.251/1.251 = 1, O = 1.881/1.251 = 1.503
Multiplying both by 2 to get whole numbers: Fe = 2, O = 3
Empirical formula = **Fe_2O_3**

Question 1.4

Question: Calculate the amount of carbon dioxide that could be produced when (i) 1 mole of carbon is burnt in air. (ii) 1 mole of carbon is burnt in 16 g of dioxygen. (iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Solution: Reaction: $C(s) + O_2(g) \rightarrow CO_2(g)$
(i) In excess air, O_2 is not limiting, so 1 mole C gives 1 mole CO_2 = **44.0 g CO_2.**
(ii) 16 g O_2 = 16/32.00 = 0.5 mol O_2. Since 1 mol C requires 1 mol O_2 but only 0.5 mol O_2 is available, O_2 is the limiting reagent. Moles of CO_2 formed = 0.5 mol = **22.0 g CO_2.**
(iii) With 2 mol C and 0.5 mol O_2 (16 g), O_2 is again limiting (needs 2 mol O_2 for 2 mol C). CO_2 formed = 0.5 mol = **22.0 g CO_2.**

Question 1.5

Question: Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}.

Solution: Moles required = Molarity × Volume (in L) = 0.375 mol L^{-1} × 0.500 L = 0.1875 mol
Mass = moles × molar mass = 0.1875 × 82.0245 = **15.38 g**

Question 1.6

Question: Calculate the concentration of nitric acid in moles per litre in a sample which has a density of 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

Solution: Consider 1000 mL (1 L) of solution.
Mass of solution = 1000 × 1.41 = 1410 g
Mass of HNO_3 = 69% of 1410 = 972.9 g
Molar mass of HNO_3 = 1.008 + 14.007 + 3(16.00) = 63.015 g mol^{-1}
Moles of HNO_3 = 972.9 / 63.015 = 15.44 mol
Since this is present in 1 L of solution, concentration = **15.44 mol L^{-1}**

Question 1.7

Question: How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

Solution: Molar mass of $CuSO_4$ = 63.55 + 32.06 + 4(16.00) = 159.61 g mol^{-1}
Moles of $CuSO_4$ = 100 / 159.61 = 0.6265 mol
Since 1 mol $CuSO_4$ contains 1 mol Cu, moles of Cu = 0.6265 mol
Mass of Cu = 0.6265 × 63.55 = **39.81 g**

Question 1.8

Question: Determine the molecular formula of an oxide of iron, in which the mass per cent of iron and oxygen are 69.9 and 30.1, respectively. Given that the molar mass of the oxide is 159.69 g mol^{-1}.

Solution: As found in Q1.3, the empirical formula is Fe_2O_3.
Empirical formula mass = 2(55.85) + 3(16.00) = 111.70 + 48.00 = 159.70 g mol^{-1}
n = molar mass / empirical formula mass = 159.69 / 159.70 ≈ 1
Since n = 1, the molecular formula is the same as the empirical formula: **Fe_2O_3**

Question 1.9

Question: Calculate the average atomic mass of chlorine using the following data:

Isotope: ^{35}Cl, % Natural Abundance: 75.77, Molar Mass: 34.9689
Isotope: ^{37}Cl, % Natural Abundance: 24.23, Molar Mass: 36.9659

Solution: Average atomic mass = (0.7577 × 34.9689) + (0.2423 × 36.9659)
= 26.496 + 8.957 = **35.45 u** (approximately)

Question 1.10

Question: In three moles of ethane (C_2H_6), calculate the following: (i) Number of moles of carbon atoms. (ii) Number of moles of hydrogen atoms. (iii) Number of molecules of ethane.

Solution:
(i) Each mole of C_2H_6 has 2 mol of C atoms, so 3 mol ethane gives 3 × 2 = **6 mol carbon atoms**
(ii) Each mole of C_2H_6 has 6 mol of H atoms, so 3 mol ethane gives 3 × 6 = **18 mol hydrogen atoms**
(iii) Number of molecules = 3 × 6.022 × 10^{23} = **1.807 × 10^{24} molecules**

Question 1.11

Question: What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

Solution: Molar mass of $C_{12}H_{22}O_{11}$ = 12(12.011) + 22(1.008) + 11(16.00) = 144.13 + 22.18 + 176.00 = 342.31 g mol^{-1}
Moles of sugar = 20 / 342.31 = 0.05843 mol
Concentration = 0.05843 mol / 2 L = **0.0292 mol L^{-1}**

Question 1.12

Question: If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution?

Solution: Moles of methanol needed = 0.25 mol L^{-1} × 2.5 L = 0.625 mol
Molar mass of CH_3OH = 12.011 + 4(1.008) + 16.00 = 32.04 g mol^{-1}
Mass of methanol needed = 0.625 × 32.04 = 20.03 g
Density = 0.793 kg L^{-1} = 793 g L^{-1}
Volume = mass / density = 20.03 / 793 = 0.02525 L = **25.3 mL**

Question 1.13

Question: Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal, is as shown below: 1 Pa = 1 N m^{-2}. If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

Solution: Mass per unit area = 1034 g cm^{-2} = 1.034 kg cm^{-2}
Converting cm^2 to m^2 (1 cm^2 = 10^{-4} m^2): mass per m^2 = 1.034 / 10^{-4} = 10340 kg m^{-2}
Force = mass × g = 10340 × 9.8 = 101332 N
Pressure = Force / Area = 101332 N / 1 m^2 ≈ **1.01 × 10^5 Pa**

Question 1.14

Question: What is the SI unit of mass? How is it defined?

Solution: The SI unit of mass is the **kilogram (kg)**. Since 20 May 2019, it is defined by fixing the numerical value of the Planck constant, h, to be exactly 6.62607015 × 10^{-34} J s (kg m^2 s^{-1}), combined with the definitions of the metre and the second. (Before this redefinition, the kilogram was defined as the mass of the International Prototype of the Kilogram, a platinum–iridium cylinder kept at the BIPM, Sèvres, France.)

Question 1.15

Question: Match the following prefixes with their multiples:

Prefixes: (i) micro (ii) deca (iii) mega (iv) giga (v) femto
Multiples (as given, scrambled): 10^6, 10^9, 10^{-6}, 10^{-15}, 10 (i.e., 10^1)

Solution: Matching each prefix with its correct SI multiple:
(i) micro = $\mathbf{10^{-6}}$
(ii) deca = $\mathbf{10^{1}}$ (i.e., 10)
(iii) mega = $\mathbf{10^{6}}$
(iv) giga = $\mathbf{10^{9}}$
(v) femto = $\mathbf{10^{-15}}$

Question 1.16

Question: What do you mean by significant figures?

Solution: Significant figures are the meaningful digits present in a measured (or calculated) quantity that carry information about its precision. They include all digits that are known with certainty plus one final digit that is estimated/uncertain. A larger number of significant figures indicates a more precise measurement.

Question 1.17

Question: A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass). (i) Express this in per cent by mass. (ii) Determine the molality of chloroform in the water sample.

Solution:
(i) 15 ppm = 15 parts per 10^6 parts, so mass % = $(15 / 10^6) \times 100$ = $\mathbf{1.5 \times 10^{-3}\%}$
(ii) Consider 10^6 g of water (≈ 1000 kg), containing 15 g of $CHCl_3$.
Molar mass of $CHCl_3$ = 12.011 + 1.008 + 3(35.453) = 119.38 g mol^{-1}
Moles of $CHCl_3$ = 15 / 119.38 = 0.1256 mol
Molality = 0.1256 mol / 1000 kg = $\mathbf{1.256 \times 10^{-4}}$ **mol** $\mathbf{kg^{-1}}$

Question 1.18

Question: Express the following in scientific notation: (i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012

Solution:
(i) 0.0048 = $\mathbf{4.8 \times 10^{-3}}$
(ii) 234,000 = $\mathbf{2.34 \times 10^{5}}$
(iii) 8008 = $\mathbf{8.008 \times 10^{3}}$
(iv) 500.0 = $\mathbf{5.000 \times 10^{2}}$
(v) 6.0012 = $\mathbf{6.0012 \times 10^{0}}$

Question 1.19

Question: How many significant figures are present in the following? (i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034

Solution:
(i) 0.0025 → **2** significant figures
(ii) 208 → **3** significant figures
(iii) 5005 → **4** significant figures
(iv) 126,000 → **3** significant figures (trailing zeros without a decimal point/without scientific notation are not counted as significant)
(v) 500.0 → **4** significant figures (trailing zero after decimal point is significant)
(vi) 2.0034 → **5** significant figures

Question 1.20

Question: Round up the following up to three significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808

Solution:
(i) 34.216 → **34.2**
(ii) 10.4107 → **10.4**
(iii) 0.04597 → **0.0460** (i.e., 4.60×10^{-2})
(iv) 2808 → **2810** (i.e., 2.81×10^{3})

Question 1.21

Question: The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

(i) Mass of dinitrogen: 14 g, Mass of dioxygen: 16 g
(ii) Mass of dinitrogen: 14 g, Mass of dioxygen: 32 g
(iii) Mass of dinitrogen: 28 g, Mass of dioxygen: 32 g
(iv) Mass of dinitrogen: 28 g, Mass of dioxygen: 80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.
(b) Fill in the blanks in the following conversions: (i) 1 km = ... mm = ... pm (ii) 1 mg = ... kg = ... ng (iii) 1 mL = ... L = ... dm^3

Solution:
(a) For a fixed mass of dinitrogen (28 g, rows iii and iv), the masses of dioxygen that combine are 32 g and 80 g, which are in the simple whole-number ratio 32:80 = 2:5. This confirms the **Law of Multiple Proportions**, which states that when two elements combine to form more than one compound, the different masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers.
(b)
(i) 1 km = **10^6 mm** = **10^{15} pm**
(ii) 1 mg = **10^{-6} kg** = **10^6 ng**
(iii) 1 mL = **10^{-3} L** = **10^{-3} dm^3**

Question 1.22

Question: If the speed of light is 3.0×10^8 m s^{-1}, calculate the distance covered by light in 2.00 ns.

Solution: Distance = speed × time = 3.0×10^8 m s^{-1} × 2.00×10^{-9} s = **0.60 m**

Question 1.23

Question: In a reaction $A + B_2 \rightarrow AB_2$, identify the limiting reagent, if any, in the following reaction mixtures: (i) 300 atoms of A + 200 molecules of B_2 (ii) 2 mol A + 3 mol B_2 (iii) 100 atoms of A + 100 molecules of B_2 (iv) 5 mol A + 2.5 mol B_2 (v) 2.5 mol A + 5 mol B_2

Solution: The reaction requires A and B_2 in a 1:1 ratio.
(i) 300 atoms A need 300 molecules B_2, but only 200 are available → **B_2 is limiting**
(ii) 2 mol A needs 2 mol B_2; 3 mol B_2 is available (excess) → **A is limiting**

(iii) 100 atoms A and 100 molecules B_2 match exactly → **no limiting reagent; both are used up completely**
(iv) 5 mol A needs 5 mol B_2, only 2.5 mol available → **B_2 is limiting**
(v) 2.5 mol A needs 2.5 mol B_2; 5 mol B_2 is available (excess) → **A is limiting**

Question 1.24

Question: Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation: $N_2(g) + H_2(g) \rightarrow 2NH_3(g)$. (i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen. (ii) Will any of the two reactants remain unreacted? (iii) If yes, which one and what would be its mass?

Solution: The balanced equation is $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$.
Moles of $N_2 = 2.00 \times 10^3 / 28.02 = 71.38$ mol
Moles of $H_2 = 1.00 \times 10^3 / 2.016 = 496.03$ mol
Required H_2 for all N_2 to react = 3 × 71.38 = 214.14 mol, which is less than the 496.03 mol available. So **N_2 is the limiting reagent**.
(i) Moles of NH_3 formed = 2 × 71.38 = 142.76 mol
Mass of NH_3 = 142.76 × 17.03 = **2431 g (≈ 2.43×10^3 g)**
(ii) **Yes**, H_2 remains unreacted (it is in excess).
(iii) Unreacted H_2 = 496.03 – 214.14 = 281.89 mol
Mass of unreacted H_2 = 281.89 × 2.016 = **568.3 g**

Question 1.25

Question: How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Solution: "0.50 mol Na_2CO_3" refers only to the **amount of substance** — 0.50 mole of Na_2CO_3 — without any reference to the volume in which it is dissolved. "0.50 M Na_2CO_3" refers to **molar concentration**, meaning 0.50 mole of Na_2CO_3 is dissolved in exactly 1 litre of solution; it describes both amount and volume together.

Question 1.26

Question: If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Solution: Reaction: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. By Gay-Lussac's Law of Gaseous Volumes, gas volumes react in the same simple ratio as their coefficients (2:1:2). 10 volumes of H_2 react completely with 5 volumes of O_2 (matching the 2:1 ratio exactly), producing **10 volumes of water vapour**.

Question 1.27

Question: Convert the following into basic units: (i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

Solution:
(i) 28.7 pm = 28.7×10^{-12} m = **2.87×10^{-11} m**
(ii) 15.15 pm = 15.15×10^{-12} m = **1.515×10^{-11} m**
(iii) 25365 mg = 25.365 g = **2.5365×10^{-2} kg**

Question 1.28

Question: Which one of the following will have the largest number of atoms? (i) 1 g Au(s) (ii) 1 g Na(s) (iii) 1 g Li(s) (iv) 1 g of Cl_2(g)

Solution: Number of atoms = (mass / atomic or molar mass) × N_A
(i) Au (197.0 g mol^{-1}): 1/197.0 × 6.022 × 10^{23} = 3.06 × 10^{21} atoms
(ii) Na (22.99 g mol^{-1}): 1/22.99 × 6.022 × 10^{23} = 2.62 × 10^{22} atoms
(iii) Li (6.94 g mol^{-1}): 1/6.94 × 6.022 × 10^{23} = 8.68 × 10^{22} atoms
(iv) Cl_2 (70.90 g mol^{-1}, 2 atoms/molecule): (1/70.90) × 2 × 6.022 × 10^{23} = 1.70 × 10^{22} atoms
Since Li has the smallest atomic mass among these, 1 g of Li contains the most atoms. **(iii) 1 g Li(s) has the largest number of atoms.**

Question 1.29

Question: Calculate the molarity of a solution of ethanol in water, in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

Solution: Let moles of ethanol = 1; since mole fraction = 0.040, moles of water = (1 − 0.040)/0.040 = 24 mol.
Mass of water = 24 × 18.02 = 432.48 g
Mass of ethanol = 1 × 46.07 = 46.07 g
Total mass of solution = 432.48 + 46.07 = 478.55 g
Assuming solution density ≈ 1 g mL^{-1} (as water), volume of solution = 478.55 mL = 0.4786 L
Molarity = 1 mol / 0.4786 L = **2.09 mol L^{-1}**

Question 1.30

Question: What will be the mass of one ^{12}C atom in g?

Solution: Mass of one ^{12}C atom = molar mass / Avogadro's number = 12 g mol^{-1} / 6.022 × 10^{23} mol^{-1} = **1.993 × 10^{-23} g**

Question 1.31

Question: How many significant figures should be present in the answer of the following calculations? (i) 0.02856 × 298.15 × 0.112 / 0.5785 (ii) 5 × 5.364 (iii) 0.0125 + 0.7864 + 0.0215

Solution:
(i) The least number of significant figures among the multiplicands/divisor (0.02856 has 4, 298.15 has 5, 0.112 has 3, 0.5785 has 4) is **3**, from 0.112. Computed value = 1.6489, rounded to 3 significant figures = **1.65**.
(ii) Here "5" is treated as an exact number (not a measured quantity), so the answer's precision is set by 5.364 (4 significant figures). 5 × 5.364 = 26.82, so the answer has **4** significant figures: **26.82**.
(iii) For addition, the result is rounded to the least number of decimal places among the terms. All three terms (0.0125, 0.7864, 0.0215) have 4 decimal places, so the sum, 0.8204, is reported with **4 decimal places**: **0.8204**.

Question 1.32

Question: Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope ^{36}Ar: Isotopic molar mass 35.96755 g mol^{-1}, Abundance 0.337%
Isotope ^{38}Ar: Isotopic molar mass 37.96272 g mol^{-1}, Abundance 0.063%
Isotope ^{40}Ar: Isotopic molar mass 39.9624 g mol^{-1}, Abundance 99.600%

Solution: Molar mass = Σ(fractional abundance × isotopic molar mass)
= (0.00337 × 35.96755) + (0.00063 × 37.96272) + (0.99600 × 39.9624)
= 0.1212 + 0.0239 + 39.8025 = **39.95 g mol^{-1}**

Question 1.33

Question: Calculate the number of atoms in each of the following: (i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He

Solution:
(i) Atoms in 52 mol Ar = 52 × 6.022 × 10^{23} = **3.13 × 10^{25} atoms**
(ii) Since 1 atom of He has a mass of ≈ 4.00 u, the number of He atoms whose total mass is 52 u = 52/4.00 ≈ **13 atoms**
(iii) Moles of He in 52 g = 52/4.00 = 13.0 mol; number of atoms = 13.0 × 6.022 × 10^{23} = **7.83 × 10^{24} atoms**

Question 1.34

Question: A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Solution:
Mass of C = (12.011/44.01) × 3.38 = 0.922 g → moles C = 0.922/12.011 = 0.0768 mol
Mass of H = (2.016/18.02) × 0.690 = 0.0772 g → moles H = 0.0772/1.008 = 0.0766 mol
Ratio C : H = 0.0768 : 0.0766 ≈ 1 : 1
(i) **Empirical formula = CH** (empirical formula mass = 12.011 + 1.008 = 13.02 g mol^{-1})
(ii) At STP, molar volume = 22.4 L mol^{-1}. Moles of gas in 10.0 L = 10.0/22.4 = 0.4464 mol
Molar mass = 11.6/0.4464 = **26.0 g mol^{-1}**
(iii) n = molar mass/empirical formula mass = 26.0/13.02 ≈ 2
Molecular formula = C_2H_2 (acetylene), calculated molar mass = 2(12.011) + 2(1.008) = 26.04 g mol^{-1}, which matches.

Question 1.35

Question: Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction: $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$. What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Solution: Moles of HCl = 0.75 mol L^{-1} × 0.025 L = 0.01875 mol
From the equation, 2 mol HCl reacts with 1 mol $CaCO_3$, so moles of $CaCO_3$ needed = 0.01875/2 = 0.009375 mol
Molar mass of $CaCO_3$ = 40.08 + 12.011 + 3(16.00) = 100.09 g mol^{-1}
Mass of $CaCO_3$ = 0.009375 × 100.09 = **0.938 g**

Question 1.36

Question: Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction: $4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$. How many grams of HCl react with 5.0 g of manganese dioxide?

Solution: Molar mass of MnO_2 = 54.94 + 2(16.00) = 86.94 g mol^{-1}
Moles of MnO_2 = 5.0/86.94 = 0.05752 mol
From the equation, 4 mol HCl reacts with 1 mol MnO_2, so moles of HCl = 4 × 0.05752 = 0.2301 mol
Molar mass of HCl = 1.008 + 35.453 = 36.461 g mol^{-1}
Mass of HCl = 0.2301 × 36.461 = **8.39 g**

Notes and Extra Questions

This chapter establishes the quantitative backbone of chemistry: the mole concept, Avogadro's number, molar mass, and the laws of chemical combination (conservation of mass, definite proportions, multiple proportions, and gaseous volumes) that together justify Dalton's atomic theory. Mastering empirical/molecular formula determination, stoichiometric calculations with limiting reagents, and concentration terms (molarity, molality, mole fraction, mass percentage, ppm) here is essential, since nearly every numerical problem in Class 11 and 12 Chemistry — from equilibrium to electrochemistry — builds directly on these skills. Pay special attention to significant figures and unit conversions, as these recur throughout the syllabus and in competitive exams like JEE and NEET.