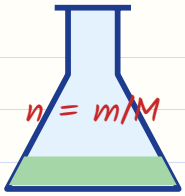
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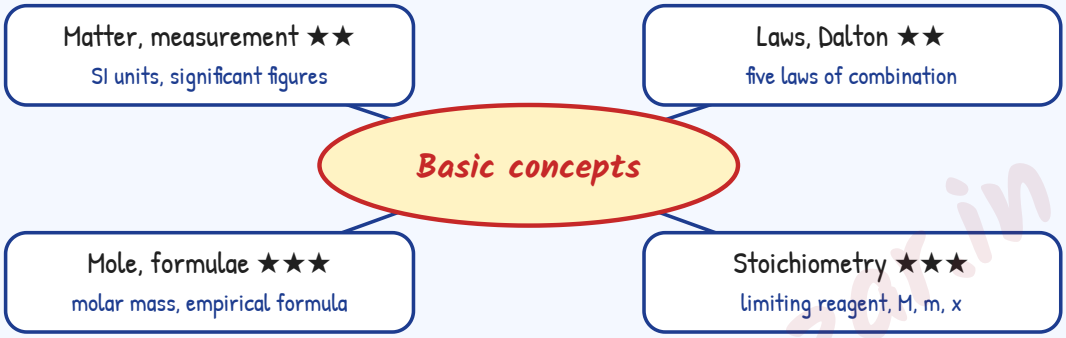


# Unit 1
# Some Basic Concepts of Chemistry




NCERT Chemistry Part I, Reprint 2026-27

## Unit map (★ = very often asked in school exams)




## What is inside

1. Concept notes with all NCERT problems solved
2. NCERT exercises (all 36), fully solved
3. Important exam questions, fully solved
4. Quick-reference sheet; practice A–G, 20 MCQs, 6 A–R, answers, summary

## Colour code used

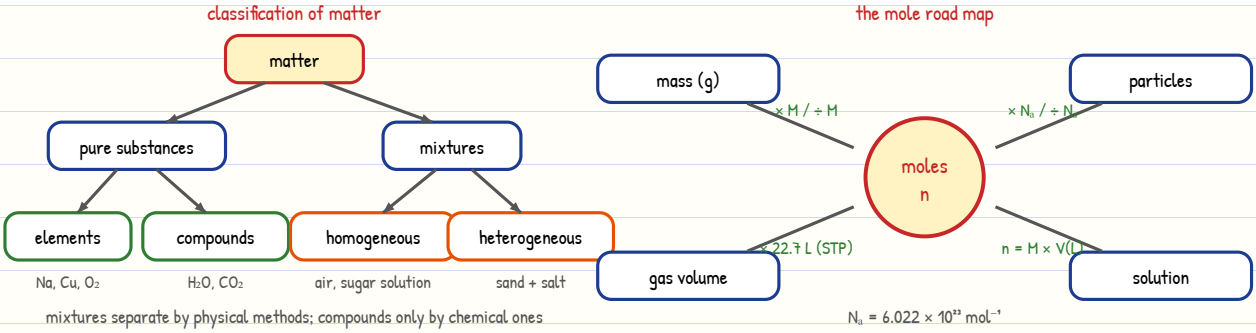
red = key term  green = definition
purple = sub-heading  black = explanation
★ = high-frequency exam topic
1 = mark the examiner gives for that point

★ **Exam tip:** Significant figures and rounding, the laws of chemical combination, the mole concept, empirical and molecular formulae, the limiting reagent, and the concentration terms (mass per cent, mole fraction, molarity, molality) are asked again and again.

NB **Prepared by Notes Bazar** · Every answer written and checked by hand against NCERT and the CBSE marking schemes. Doubt or mistake? WhatsApp 88240 98091.

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# 1.1–1.2 Chemistry and matter; 1.3–1.4 measurement; 1.5 laws; 1.6 Dalton ★★★




*Left: how matter is classified. Right: every stoichiometry problem passes through the number of moles*

## Matter

**Matter** has mass and occupies space. **States**: solid (fixed shape and volume), liquid (fixed volume), gas (neither); interconvertible by changing T and P. **Mixtures** (variable composition; components keep their properties; separable by physical methods): **homogeneous** (air, sugar solution) or **heterogeneous** (salt and sugar). **Pure substances** (fixed composition): **elements** (one kind of atom: Na, Cu, $H_2$, $O_2$) and **compounds** (two or more elements in a fixed ratio: $H_2O$, $CO_2$; separable only by chemical methods; their properties differ from those of the elements).

## Properties and measurement

**Physical properties** are measured without changing identity (colour, odour, melting point, density); **chemical properties** involve a change (acidity, combustibility). **SI base units**: metre, kilogram, second, ampere, kelvin, mole, candela. Volume: $m^3$; 1 L = 1 $dm^3$ = 1000 $cm^3$ (mL). Density = mass/volume (kg $m^{-3}$ or g $cm^{-3}$). Temperature: $K = {}^\circ C + 273.15$; ${}^\circ F = \frac{9}{5}({}^\circ C) + 32$.

## 1.4 Uncertainty in measurement

**Scientific notation** $N \times 10^n$ ($1 \le N < 10$): $0.00016 = 1.6 \times 10^{-4}$. **Precision**: agreement of repeated values; **accuracy**: agreement with the true value.

**Significant figures** (certain digits + one uncertain): non-zero digits count; zeros before the first non-zero digit do not (0.03: one); zeros between non-zero digits do (2.005: four); terminal zeros count if there is a decimal point (0.200: three); exact numbers (2 balls, 20 eggs) have infinitely many. In **addition/subtraction** the result has as many **decimal places** as the least; in **multiplication/division**, as many **significant figures** as the least. **Rounding**: if the next digit is > 5 raise the last, if < 5 leave it; if it is 5, raise only an odd last digit (6.35 → 6.4, 6.25 → 6.2). **Dimensional analysis** (factor-label method): multiply by unit factors equal to 1, e.g. 3 inch × $\frac{2.54\ \text{cm}}{1\ \text{inch}}$ = 7.62 cm.

## 1.5 Laws of chemical combination

(1) **Conservation of mass** (Lavoisier): matter can neither be created nor destroyed (total mass of products = total mass of reactants). (2) **Definite proportions** (Proust): a given compound always contains exactly the same proportion of elements by weight. (3) **Multiple proportions** (Dalton): if two elements form more than one compound, the masses of one element that combine with a fixed mass of the other are in a ratio of small whole numbers ($H_2$ 2 g with $O_2$ 16 g gives 18 g water; with 32 g gives 34 g $H_2O_2$: 16 : 32 = 1 : 2). (4) **Gay Lussac's law of gaseous volumes** (1808): gases combine or are produced in a simple ratio by volume (same T, P): 100 mL $H_2$ + 50 mL $O_2$ → 100 mL water vapour (2 : 1 : 2). (5) **Avogadro's law** (1811): equal volumes of all gases at the same T and P contain equal numbers of molecules (he distinguished atoms from molecules).

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### 1.6 Dalton's atomic theory (1808)

(i) Matter consists of indivisible **atoms**. (ii) All atoms of an element have identical properties, including mass; atoms of different elements differ in mass. (iii) Compounds form when atoms of different elements combine in a fixed ratio. (iv) Chemical reactions involve **reorganisation of atoms**; atoms are neither created nor destroyed. It explained the laws of combination, but not the laws of gaseous volumes or why atoms combine (and atoms were later found to be divisible, with isotopes).

### 1.7 Atomic and molecular masses

Atomic masses are relative to **carbon-12**, assigned exactly 12 **atomic mass units**; 1 amu (now **u**, unified mass) = 1/12 of the mass of a $^{12}C$ atom = $1.66056 \times 10^{-24}$ g. H = 1.0080 u, O = 15.995 u (mass spectrometry). **Average atomic mass** from isotopic abundances: carbon ($^{12}C$ 98.892%, $^{13}C$ 1.108%, $^{14}C$ $2 \times 10^{-10}$%) = 12.011 u. **Molecular mass** = sum of atomic masses: $CH_4$ = 12.011 + 4(1.008) = 16.043 u; $H_2O$ = 18.02 u. **Formula mass** for ionic solids without discrete molecules: NaCl = 23.0 + 35.5 = 58.5 u.

**Problem 1.1** Glucose $C_6H_{12}O_6$: 6(12.011) + 12(1.008) + 6(16.00) = 72.066 + 12.096 + 96.00 = **180.162 u**.

## 1.8 Mole; 1.9 percentage composition; 1.10 stoichiometry and concentration ★★★

### 1.8 Mole concept and molar mass

The **mole** (mol) is the SI unit of amount of substance: one mole contains exactly **$6.02214076 \times 10^{23}$** elementary entities (atoms, molecules, ions, electrons…), the **Avogadro constant $N_A$**. (From the mass of a $^{12}C$ atom, $1.992648 \times 10^{-23}$ g: 12 g/$1.992648 \times 10^{-23}$ g = $6.022 \times 10^{23}$ atoms.) The **molar mass** (g mol$^{-1}$) is numerically equal to the atomic, molecular or formula mass in u: water 18.02 g mol$^{-1}$, NaCl 58.5 g mol$^{-1}$. So $n = \frac{m}{M} = \frac{N}{N_A}$; at STP (273.15 K, 1 bar) one mole of an ideal gas occupies 22.7 L (22.4 L at 1 atm).

### 1.9 Percentage composition, formulae

Mass % of an element = $\frac{\text{mass of that element in the compound}}{\text{molar mass of the compound}} \times 100$. Ethanol $C_2H_5OH$ (46.068 g mol$^{-1}$): C 52.14%, H 13.13%, O 34.73%.

The **empirical formula** gives the simplest whole-number ratio of atoms; the **molecular formula** the exact numbers: molecular formula = n × (empirical formula), with n = molar mass/empirical formula mass.

**Problem 1.2** 4.07% H, 24.27% C, 71.65% Cl; molar mass 98.96 g. In 100 g: H 4.07/1.008 = 4.04 mol; C 24.27/12.01 = 2.021 mol; Cl 71.65/35.453 = 2.021 mol. Divide by 2.021: H : C : Cl = 2 : 1 : 1, empirical formula **$CH_2Cl$** (49.48 g). n = 98.96/49.48 = 2: molecular formula **$C_2H_4Cl_2$**.

### 1.10 Stoichiometry

A balanced equation gives the mole (and, for gases, volume) ratios:
$\mathrm{CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)}$: 1 mol $CH_4$ + 2 mol $O_2$ → 1 mol $CO_2$ + 2 mol $H_2O$; 16 g $CH_4$ + 64 g $O_2$ → 44 g $CO_2$ + 36 g $H_2O$; 22.7 L $CH_4$ + 45.4 L $O_2$ → 22.7 L $CO_2$ + 45.4 L $H_2O$(g). **Balancing** (e.g. propane): write correct formulas, balance C, then H, then O: $\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$; never change subscripts.

**Problem 1.3** Water from 16 g of methane: 1 mol $CH_4$ gives 2 mol $H_2O$ = 2 × 18 = **36 g**. **Problem 1.4** Methane for 22

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g $CO_2$: 22/44 = 0.5 mol $CO_2$, from **0.5 mol $CH_4$**.

## Limiting reagent

The reactant present in the lesser amount (relative to the equation) is used up first and limits the amount of product: the **limiting reagent**; the other is in excess.

**Problem 1.5** 50.0 kg $N_2$ + 10.0 kg $H_2$ for $\mathrm{N_2 + 3\,H_2 \longrightarrow 2\,NH_3}$. n($N_2$) = 50000/28.0 = 1.786 × $10^3$ mol; n($H_2$) = 10000/2.016 = 4.96 × $10^3$ mol. $N_2$ would need 5.36 × $10^3$ mol $H_2$, more than available: **$H_2$ is limiting**. $NH_3$ = 4.96 × $10^3$ × 2/3 = 3.30 × $10^3$ mol × 17.0 = **56.1 kg**.

## Reactions in solution: concentration

(1) **Mass per cent** = $\dfrac{\text{mass of solute}}{\text{mass of solution}} \times 100$. (2) **Mole fraction** $x_A = \dfrac{n_A}{n_A + n_B}$ ($x_A + x_B = 1$). (3) **Molarity** M = moles of solute per litre of solution (depends on temperature); dilution: $M_1V_1 = M_2V_2$ (200 mL of 1 M NaOH diluted to 1 L gives 0.2 M). (4) **Molality** m = moles of solute per kg of solvent (independent of temperature).

**Problem 1.6** 2 g A in 18 g water: mass % = 2/20 × 100 = **10%**. **Problem 1.7** 4 g NaOH in 250 mL: n = 4/40 = 0.1 mol; M = 0.1/0.250 = **0.4 M**. **Problem 1.8** 3 M NaCl, density 1.25 g $mL^{-1}$: in 1 L, NaCl = 3 × 58.5 = 175.5 g, solution 1250 g, water 1074.5 g: molality = 3/1.0745 = **2.79 m**.

## Summary (NCERT)

- Matter: elements, compounds, mixtures; physical and chemical properties; SI units.
- Scientific notation and significant figures express the uncertainty of measurements.
- Five laws of chemical combination; Dalton's atomic theory.
- Atomic mass relative to $^{12}C$; average atomic mass; molecular and formula mass.
- Mole: 6.022 × $10^{23}$ entities; molar mass; empirical and molecular formulae.
- Stoichiometry from balanced equations; limiting reagent; mass %, mole fraction, molarity, molality.

✗ **Don't do this:** Molarity uses litres of **solution**; molality uses kilograms of **solvent**.

## NCERT exercises, fully solved

*(H 1.008, C 12.01, N 14.01, O 16.00, Na 22.99, S 32.06, Cl 35.45, Ca 40.08, Mn 54.94, Fe 55.85, Cu 63.55)*

**1.1** Molar mass of (i) $H_2O$ (ii) $CO_2$ (iii) $CH_4$.

> **Ans.** (i) 2(1.008) + 16.00 = **18.02 g $mol^{-1}$**. (ii) 12.01 + 2(16.00) = **44.01 g $mol^{-1}$**. (iii) 12.01 + 4(1.008) = **16.04 g $mol^{-1}$**.

**1.2** Mass per cent of the elements in $Na_2SO_4$.

> **Ans.** M = 2(22.99) + 32.06 + 4(16.00) = 142.04 g $mol^{-1}$. Na = 45.98/142.04 × 100 = **32.37%**; S = 32.06/142.04 × 100 = **22.57%**; O = 64.00/142.04 × 100 = **45.06%**.

**1.3** Empirical formula of an iron oxide with 69.9% Fe and 30.1% O.

> **Ans.** Fe: 69.9/55.85 = 1.25 mol; O: 30.1/16.00 = 1.88 mol. Ratio 1 : 1.5 = 2 : 3: **$Fe_2O_3$**.

**1.4** $CO_2$ produced when (i) 1 mol C burns in air (ii) 1 mol C burns in 16 g $O_2$ (iii) 2 mol C burn in 16 g $O_2$.

> **Ans.** $\mathrm{C + O_2 \longrightarrow CO_2}$. (i) Air is in excess: 1 mol $CO_2$ = **44 g**. (ii) 16 g $O_2$ = 0.5 mol, the limiting reagent: 0.5 mol $CO_2$ = **22 g**. (iii) Again 0.5 mol $O_2$ limits: **22 g**.

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**1.5** Mass of sodium acetate (82.0245 g $mol^{-1}$) for 500 mL of a 0.375 M solution.

> **Ans.** n = 0.375 × 0.500 = 0.1875 mol; mass = 0.1875 × 82.0245 = **15.38 g**.

**1.6** Molarity of nitric acid of density 1.41 g $mL^{-1}$ and 69% by mass.

> **Ans.** 1 L = 1410 g of solution containing 0.69 × 1410 = 972.9 g $HNO_3$ (63.02 g $mol^{-1}$): 972.9/63.02 = **15.44 M**.

**1.7** How much copper can be obtained from 100 g of $CuSO_4$?

> **Ans.** $M(CuSO_4)$ = 63.55 + 32.06 + 64.00 = 159.61 g $mol^{-1}$; one mole contains 63.55 g Cu. From 100 g: 100 × 63.55/159.61 = **39.8 g**.

**1.8** Molecular formula of an iron oxide with 69.9% Fe and 30.1% O.

> **Ans.** Empirical formula $Fe_2O_3$ (Exercise 1.3), empirical formula mass 159.7 g $mol^{-1}$. The molar mass of iron(III) oxide is also 159.7 g $mol^{-1}$, so n = 1: **$Fe_2O_3$**.

**1.9** Average atomic mass of chlorine: $^{35}Cl$ 75.77% (34.9689), $^{37}Cl$ 24.23% (36.9659).

> **Ans.** 0.7577 × 34.9689 + 0.2423 × 36.9659 = 26.4959 + 8.9568 = **35.45 u**.

**1.10** In 3 mol of ethane $C_2H_6$: (i) moles of C atoms (ii) moles of H atoms (iii) number of molecules.

> **Ans.** (i) 3 × 2 = **6 mol**. (ii) 3 × 6 = **18 mol**. (iii) 3 × 6.022 × $10^{23}$ = **1.807 × $10^{24}$**.

**1.11** Concentration of sugar $C_{12}H_{22}O_{11}$ if 20 g are dissolved to make 2 L.

> **Ans.** M = 342.3 g $mol^{-1}$; n = 20/342.3 = 0.0584 mol; concentration = 0.0584/2 = **0.0292 mol $L^{-1}$**.

**1.12** Volume of methanol (density 0.793 kg $L^{-1}$) for 2.5 L of 0.25 M solution.

> **Ans.** n = 0.25 × 2.5 = 0.625 mol × 32.04 = 20.03 g; volume = 20.03/0.793 = **25.3 mL**.

**1.13** Mass of air at sea level is 1034 g $cm^{-2}$. Pressure in pascal?

> **Ans.** Force on 1 $cm^2$ = 1.034 kg × 9.8 m $s^{-2}$ = 10.13 N; area $10^{-4}$ $m^2$: P = 10.13/$10^{-4}$ = **1.01 × $10^5$ Pa**.

**1.14** What is the SI unit of mass? How is it defined?

> **Ans.** The **kilogram** (kg). It is defined by fixing the numerical value of the Planck constant h at 6.626 07015 × $10^{-34}$ J s (= kg $m^2$ $s^{-1}$), the metre and second being defined through c and the caesium frequency. (Earlier it was the mass of the international prototype kilogram.)

**1.15** Match the prefixes micro, deca, mega, giga, femto with $10^6$, $10^9$, $10^{-6}$, $10^{-15}$, 10.

> **Ans.** micro → **$10^{-6}$**; deca → **10**; mega → **$10^6$**; giga → **$10^9$**; femto → **$10^{-15}$**.

**1.16** What do you mean by significant figures?

> **Ans.** The meaningful digits in a measured quantity: all the digits known with certainty plus one uncertain digit. They show the precision of a measurement (e.g. 11.2 mL has three).

**1.17** Drinking water contains 15 ppm (by mass) of chloroform $CHCl_3$. (i) Express in mass per cent. (ii) Molality?

> **Ans.** (i) 15 g per $10^6$ g: 15/$10^6$ × 100 = **1.5 × $10^{-3}$ %** (15 × $10^{-4}$ %). (ii) $M(CHCl_3)$ = 119.4 g $mol^{-1}$; 15 g = 0.1257 mol in about $10^6$ g = $10^3$ kg water: molality ≈ **1.26 × $10^{-4}$ m**.

**1.18** Express in scientific notation: (i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012.

> **Ans.** (i) **4.8 × $10^{-3}$** (ii) **2.34 × $10^5$** (iii) **8.008 × $10^3$** (iv) **5.000 × $10^2$** (v) **6.0012 × $10^0$**.

**1.19** Significant figures in (i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034.

> **Ans.** (i) **2** (ii) **3** (iii) **4** (iv) **3** (v) **4** (vi) **5**.

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**1.20** Round to three significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808.

> **Ans.** (i) **34.2** (ii) **10.4** (iii) **0.0460** (iv) **2810**.

**1.21** $N_2$ and $O_2$ combine as 14 g : 16 g, 14 : 32, 28 : 32, 28 : 80. (a) Which law is obeyed? State it. (b) Fill in: 1 km = __ mm = __ pm; 1 mg = __ kg = __ ng; 1 mL = __ L = __ $dm^3$.

> **Ans.** (a) For a fixed 28 g of $N_2$, the $O_2$ masses are 32, 64, 32, 80 g, i.e. 2 : 4 : 2 : 5 (small whole numbers): the **law of multiple proportions**: when two elements form more than one compound, the masses of one that combine with a fixed mass of the other are in a simple whole-number ratio. (b) 1 km = $10^6$ mm = $10^{15}$ pm; 1 mg = $10^{-6}$ kg = $10^6$ ng; 1 mL = $10^{-3}$ L = $10^{-3}$ $dm^3$.

**1.22** Distance light ($3.0 \times 10^8$ m $s^{-1}$) covers in 2.00 ns.

> **Ans.** $3.0 \times 10^8 \times 2.00 \times 10^{-9}$ = **0.600 m**.

**1.23** For $A + B_2 \rightarrow AB_2$, find the limiting reagent: (i) 300 atoms A + 200 molecules B (ii) 2 mol A + 3 mol B (iii) 100 atoms A + 100 molecules B (iv) 5 mol A + 2.5 mol B (v) 2.5 mol A + 5 mol B.

> **Ans.** A and $B_2$ react 1 : 1. (i) **B** limiting (200 < 300). (ii) **A** limiting. (iii) Stoichiometric mixture: **none**. (iv) **B** limiting. (v) **A** limiting.

**1.24** $N_2 + 3\,H_2 \longrightarrow 2\,NH_3$ with $2.00 \times 10^3$ g $N_2$ and $1.00 \times 10^3$ g $H_2$. (i) Mass of $NH_3$? (ii) Is any reactant left? (iii) Which, and how much?

> **Ans.** $n(N_2)$ = 2000/28.02 = 71.4 mol; $n(H_2)$ = 1000/2.016 = 496 mol; $N_2$ needs 214 mol $H_2$, so **$N_2$ is limiting**. (i) $NH_3$ = 142.8 mol × 17.03 = **$2.43 \times 10^3$ g**. (ii) **Yes**. (iii) **Hydrogen**: 496 − 214 = 282 mol = **$5.68 \times 10^2$ g** (key: $5.72 \times 10^2$ g with rounded values).

**1.25** How are 0.50 mol $Na_2CO_3$ and 0.50 M $Na_2CO_3$ different?

> **Ans.** 0.50 mol is an amount: 0.50 × 106 = **53 g** of $Na_2CO_3$. 0.50 M is a concentration: 0.50 mol (53 g) dissolved in **1 L of solution**.

**1.26** 10 volumes of $H_2$ react with 5 volumes of $O_2$. Volumes of water vapour?

> **Ans.** $2\,H_2 + O_2 \longrightarrow 2\,H_2O$: 2 : 1 : 2 by volume (Gay Lussac), so **10 volumes** of water vapour.

**1.27** Convert to basic units: (i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg.

> **Ans.** (i) **$2.87 \times 10^{-11}$ m** (ii) **$1.515 \times 10^{-11}$ m** (iii) **$2.5365 \times 10^{-2}$ kg**.

**1.28** Which has the most atoms: 1 g Au, 1 g Na, 1 g Li, 1 g $Cl_2$?

> **Ans.** Moles of atoms: Au 1/197 = 0.0051; Na 1/22.99 = 0.0435; Li 1/6.94 = 0.144; Cl 1/35.45 = 0.0282. **1 g Li** (the lightest atoms).

**1.29** Molarity of ethanol in water if its mole fraction is 0.040 (density of water 1).

> **Ans.** Take 1 L ≈ 1000 g of water = 1000/18 = 55.5 mol. With $x = \dfrac{n}{n + 55.5} = 0.040$: $n = \dfrac{0.040 \times 55.5}{0.960} = 2.31$ mol. Taking the solution volume ≈ 1 L: **≈ 2.31 M**.

**1.30** Mass of one $^{12}C$ atom in g?

> **Ans.** 12 g/$6.022 \times 10^{23}$ = **$1.99 \times 10^{-23}$ g**.

**1.31** Significant figures in the answers of (i) 0.02856 × 298.15 × 0.112/0.5785 (ii) 5 × 5.364 (iii) 0.0125 + 0.7864 + 0.0215.

> **Ans.** (i) **3** (0.112 has three). (ii) **4** (5 is exact). (iii) Sum 0.8204, four decimal places: **4**.

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**1.32** Molar mass of natural argon: $^{36}Ar$ 35.96755 (0.337%), $^{38}Ar$ 37.96272 (0.063%), $^{40}Ar$ 39.9624 (99.600%).

> **Ans.** 0.00337 × 35.96755 + 0.00063 × 37.96272 + 0.99600 × 39.9624 = 0.1212 + 0.0239 + 39.8026 = **39.948 g mol$^{-1}$**.

**1.33** Number of atoms in (i) 52 mol Ar (ii) 52 u He (iii) 52 g He.

> **Ans.** (i) 52 × 6.022 × $10^{23}$ = **3.131 × $10^{25}$**. (ii) One He atom is 4 u: 52/4 = **13 atoms**. (iii) 52/4 = 13 mol × 6.022 × $10^{23}$ = **7.83 × $10^{24}$**.

**1.34** A welding gas (C and H only) gives 3.38 g $CO_2$ and 0.690 g $H_2O$ on burning; 10.0 L at STP weigh 11.6 g. (i) Empirical formula (ii) molar mass (iii) molecular formula.

> **Ans.** C = 3.38 × 12/44 = 0.922 g (0.0768 mol); H = 0.690 × 2/18 = 0.0767 g (0.0767 mol). (i) C : H = 1 : 1: **CH**. (ii) M = 11.6 × 22.4/10.0 = **26.0 g mol$^{-1}$** (26.3 with 22.7 L). (iii) n = 26/13 = 2: **$C_2H_2$** (ethyne).

**1.35** Mass of $CaCO_3$ to react with 25 mL of 0.75 M HCl: $\mathrm{CaCO_3 + 2\,HCl \longrightarrow CaCl_2 + CO_2 + H_2O}$.

> **Ans.** n(HCl) = 0.025 × 0.75 = 0.01875 mol; $CaCO_3$ = 0.009375 mol × 100.09 = **0.94 g**.

**1.36** $\mathrm{4\,HCl + MnO_2 \longrightarrow 2\,H_2O + MnCl_2 + Cl_2}$. Mass of HCl reacting with 5.0 g $MnO_2$?

> **Ans.** n($MnO_2$) = 5.0/86.94 = 0.0575 mol; HCl = 4 × 0.0575 = 0.230 mol × 36.46 = **8.4 g**.

## Important exam questions, fully solved

Questions of the kind set in school and annual exams, with complete answers and mark steps.

### A. MCQs and Assertion–Reason (1 mark each)

*Exam-style* The number of significant figures in 0.00230 is (A) 5 (B) 2 (C) 3 (D) 6 [1]

> **Ans.** Leading zeros do not count; 2, 3 and the trailing zero after the decimal do: 3. **(C)**.

*Exam-style* The mass of one mole of electrons ($m_e$ = 9.11 × $10^{-31}$ kg) is (A) 0.55 mg (B) 0.55 kg (C) 5.5 g (D) 9.11 g [1]

> **Ans.** 9.11 × $10^{-31}$ × 6.022 × $10^{23}$ = 5.49 × $10^{-7}$ kg = 0.55 mg. **(A)**.

*Exam-style* Molality is preferred to molarity because it (A) is easier to measure (B) does not depend on temperature (C) uses moles (D) uses litres [1]

> **Ans.** Molality uses mass of solvent, which does not change with temperature. **(B)**.

*Exam-style* The law illustrated by $H_2O$ and $H_2O_2$ is (A) definite proportions (B) conservation of mass (C) gaseous volumes (D) multiple proportions [1]

> **Ans.** 16 g and 32 g O with 2 g H: 1 : 2. **(D)**.

*Exam-style* The empirical formula of glucose $C_6H_{12}O_6$ is (A) $CH_2O$ (B) $C_6H_{12}O_6$ (C) $C_2H_4O_2$ (D) CHO [1]

> **Ans.** Divide by 6: $CH_2O$. **(A)**.

*Exam-style* Assertion (A): The limiting reagent determines the amount of product. Reason (R): It is completely consumed first. [1]

> **Ans.** Both true and R explains A. **(A)**.

*Exam-style* Assertion (A): One mole of $H_2$ and one mole of $O_2$ have the same mass. Reason (R): They contain the same number of molecules. [1]

> **Ans.** A false (2 g vs 32 g); R true. **(D)**.

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## B. Short answer (2 marks)

Exam-style Distinguish between molarity and molality. [2]

**Ans.** Molarity = moles of solute per litre of solution (mol $L^{-1}$); depends on temperature (1). Molality = moles of solute per kg of solvent (mol $kg^{-1}$); independent of temperature (1).

Exam-style Calculate the number of molecules in 9 g of water. [2]

**Ans.** n = 9/18 = 0.5 mol (1); $N = 0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$ (1).

Exam-style State Avogadro's law and Gay Lussac's law of gaseous volumes. [2]

**Ans.** Avogadro: equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules (1). Gay Lussac: gases combine or are produced in a simple whole-number ratio by volume, at the same T and P (1).

Exam-style Calculate the volume of 0.1 M NaOH needed to prepare 500 mL of 0.02 M NaOH. [2]

**Ans.** $M_1V_1 = M_2V_2$ (½): $0.1 \times V_1 = 0.02 \times 500$ (½); $V_1$ = 100 mL, diluted to 500 mL (1).

Exam-style Give two limitations of Dalton's atomic theory. [2]

**Ans.** Atoms are divisible (electrons, protons, neutrons) (1); atoms of the same element can have different masses (isotopes); it could not explain Gay Lussac's law or why atoms combine (any two) (1).

## C. Short answer (3 marks)

Exam-style A compound has 40% C, 6.67% H and 53.33% O; molar mass 180 g $mol^{-1}$. Find its molecular formula. [3]

**Ans.** Moles in 100 g: C 40/12 = 3.33, H 6.67/1 = 6.67, O 53.33/16 = 3.33 (1). Ratio 1 : 2 : 1: $CH_2O$ (30 g $mol^{-1}$) (1). n = 180/30 = 6: $C_6H_{12}O_6$ (1).

Exam-style 4 g of $H_2$ and 32 g of $O_2$ react to form water. Find the limiting reagent and the mass of water. [3]

**Ans.** $2\,H_2 + O_2 \longrightarrow 2\,H_2O$; $n(H_2)$ = 2 mol, $n(O_2)$ = 1 mol: exactly 2 : 1 (1). Neither is in excess (stoichiometric) (1); water = 2 mol = 36 g (1).

Exam-style A solution has 25% (by mass) ethanol $C_2H_5OH$ in water. Find the mole fraction of ethanol. [3]

**Ans.** In 100 g: 25 g ethanol = 25/46 = 0.543 mol; 75 g water = 75/18 = 4.167 mol (1½). $x = \frac{0.543}{0.543 + 4.167} = 0.115$ (1½).

## D. Long answer (5 marks)

Exam-style (a) State the five laws of chemical combination. (b) Show that the data 2 g H + 16 g O (water) and 2 g H + 32 g O ($H_2O_2$) illustrate one of them. (c) What mass of oxygen is needed to burn 6 g of carbon completely? [5]

**Ans.** (a) Conservation of mass; definite proportions; multiple proportions; Gay Lussac's law of gaseous volumes; Avogadro's law (with statements) (2½). (b) Masses of O combining with a fixed 2 g of H: 16 : 32 = 1 : 2, a simple ratio: law of multiple proportions (1). (c) $C + O_2 \longrightarrow CO_2$: 12 g C needs 32 g $O_2$, so 6 g needs 16 g (1½).

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## E. Case study (4 marks)

**Case (exam-style):** A lab technician prepares 250 mL of 0.5 M $H_2SO_4$ from a concentrated acid of density 1.84 g $mL^{-1}$ containing 98% $H_2SO_4$ by mass (M = 98 g $mol^{-1}$).

(a) Find the molarity of the concentrated acid. [1] **Ans.** 1 L = 1840 g with 0.98 × 1840 = 1803 g $H_2SO_4$ = 18.4 mol: 18.4 M.

(b) What volume of it is needed? [1] **Ans.** 18.4 × V = 0.5 × 250 ⇒ V = 6.8 mL.

(c) Why must the acid be added to water and not water to acid? How many moles of $H_2SO_4$ are in the final solution? [2] **Ans.** Dilution is strongly exothermic; adding water to the acid can make it boil and splash. Moles = 0.5 × 0.250 = 0.125 mol.

## Quick-reference sheet

| Item | Key fact |
|---|---|
| mole | $N_A = 6.022 \times 10^{23}$ entities |
| moles | n = m/M = N/$N_A$ = V/22.7 L (gas, STP) |
| 1 u | 1/12 of a $^{12}C$ atom = $1.66 \times 10^{-24}$ g |
| mass % | mass of element/molar mass × 100 |
| molecular formula | n × empirical; n = M/(EF mass) |
| molarity | mol per L of solution; $M_1V_1 = M_2V_2$ |
| molality | mol per kg of solvent |
| mole fraction | $x_A = n_A/(n_A + n_B)$ |
| s.f. rules | +/−: decimal places; ×/÷: s.f. |

**✗ Don't do this:** Balance the equation before using mole ratios; never change subscripts.

**✗ Don't do this:** Find the limiting reagent before calculating products.

**Case study (practice): an antacid tablet.** An antacid tablet contains 0.5 g of $CaCO_3$; stomach acid is about 0.1 M HCl. ($CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$)

(i) How many moles of $CaCO_3$ are in the tablet? (ii) What volume of 0.1 M HCl does it neutralise? (iii) What volume of $CO_2$ (STP, 22.7 L $mol^{-1}$) is released? (iv) Which reactant is limiting if 50 mL of the acid is present?

# Practice set: test yourself *Exam-style*

Try every question before looking at the answers at the end of the chapter. 64 or more out of 80 means you are ready for the exam.

## A. Fill in the blanks (1 mark each)

1. The SI unit of amount of substance is the _____.
2. $N_A$ = _____ $mol^{-1}$.
3. 1 u = 1/12 of the mass of a _____ atom.
4. The law of definite proportions was given by _____.
5. Molarity depends on _____.
6. Empirical formula of benzene $C_6H_6$ is _____.

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7. Molar mass of $CO_2$ = _____.
8. Air is a _____ mixture.

## B. True or false? (1 mark each)

1. Compounds can be separated by physical methods.
2. Mass and weight are the same.
3. Exact numbers have infinite significant figures.
4. Molality changes with temperature.
5. Dalton said atoms are indivisible.
6. The sum of mole fractions is 1.
7. 6.25 rounds to 6.3 (three s.f. → two).
8. Avogadro distinguished atoms and molecules.

## C. Very short answer (1 mark each)

1. Give the molecular mass of glucose.
2. Convert 25 °C to kelvin.
3. Give the number of s.f. in 2.005.
4. Write 0.00016 in scientific notation.
5. Give the mass of 0.5 mol of NaCl.
6. State the law of conservation of mass.
7. Give an example of a heterogeneous mixture.
8. What is the formula mass of NaCl?
9. Give the molarity of 4 g NaOH in 250 mL.
10. What is a limiting reagent?

## D. Short answer I (2 marks each)

1. Find the mass per cent of C, H, O in ethanol.
2. Find the empirical formula for 4.07% H, 24.27% C, 71.65% Cl.
3. Find the water from 16 g methane.
4. Find the molality of 3 M NaCl of density 1.25 g/mL.
5. Find the mass per cent of 2 g A in 18 g water.
6. Distinguish precision and accuracy.
7. Find the average atomic mass of chlorine.
8. Give the rules for s.f. in addition and multiplication.

**My score** (fill in after checking the answers)

| Section | A /8 | B /8 | C /10 | D /16 | E /18 | F /20 | G (bonus) | Total /80 |
|---|---|---|---|---|---|---|---|---|
| My marks | | | | | | | | |

Rough work

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## E. Short answer II (3 marks each)

1. Find the limiting reagent and $NH_3$ formed from 50.0 kg $N_2$ and 10.0 kg $H_2$.
2. Find the molecular formula in Problem 1.2.
3. Explain the classification of matter with examples.
4. State Dalton's atomic theory.
5. Find the molarity of 69% $HNO_3$ of density 1.41 g/mL.
6. Balance the combustion of propane and find the $O_2$ for 1 mol.

## F. Long answer (5 marks each)

1. Explain the mole concept, molar mass and the methods of expressing concentration, with one numerical each.
2. Explain significant figures: rules, arithmetic and rounding; and dimensional analysis with examples.

## G. HOTS (higher-order thinking)

1. Why is the atomic mass of chlorine 35.45 and not a whole number?
2. Which has more molecules: 1 g of $H_2$ or 1 g of $O_2$?
3. Why is molality used for colligative properties?
4. Can the mass of products ever differ from that of reactants?

# 20 MCQs

1. A homogeneous mixture:
   (a) sand in water (b) oil in water (c) sugar solution (d) grains and pulses
2. SI unit of mass:
   (a) kilogram (b) gram (c) pound (d) tonne
3. Significant figures in 5005:
   (a) 2 (b) 1 (c) 3 (d) 4
4. 0.04597 to three s.f.:
   (a) 0.0459 (b) 0.0460 (c) 0.046 (d) 0.04600
5. $N_2$ : $O_2$ data (Ex. 1.21) obey
   (a) definite proportions (b) multiple proportions (c) conservation (d) Avogadro
6. Mass of one $^{12}C$ atom:
   (a) 12 g (b) 1 u (c) $6.02 \times 10^{23}$ g (d) $1.99 \times 10^{-23}$ g
7. Molar mass of $CH_4$:
   (a) 16.04 (b) 12.01 (c) 18.02 (d) 44.01 g/mol
8. % Na in $Na_2SO_4$:
   (a) 22.57 (b) 45.06 (c) 32.37 (d) 16.2
9. Empirical formula of 69.9% Fe oxide:
   (a) FeO (b) $FeO_2$ (c) $Fe_3O_4$ (d) $Fe_2O_3$

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10. Copper from 100 g $CuSO_4$:
(a) 63.5 g (b) 39.8 g (c) 25.4 g (d) 100 g

11. Molarity of 69% $HNO_3$ (1.41 g/mL):
(a) 6.9 M (b) 1.41 M (c) 15.44 M (d) 22.4 M

12. Sugar 20 g in 2 L:
(a) 0.0292 M (b) 0.0584 M (c) 0.29 M (d) 10 M

13. Largest number of atoms (1 g):
(a) Au (b) Li (c) Na (d) $Cl_2$

14. Limiting reagent: 2 mol A + 3 mol $B_2$ ($A + B_2 \rightarrow AB_2$):
(a) B (b) none (c) A (d) both

15. $NH_3$ from 2 kg $N_2$ and 1 kg $H_2$:
(a) 1.21 kg (b) 0.5 kg (c) 3.0 kg (d) 2.43 kg

16. Volumes of water vapour from 10 vol $H_2$ + 5 vol $O_2$:
(a) 10 (b) 5 (c) 15 (d) 20

17. Welding gas molecular formula:
(a) $C_2H_2$ (b) $CH_4$ (c) $C_2H_4$ (d) $C_6H_6$

18. $CaCO_3$ for 25 mL of 0.75 M HCl:
(a) 1.88 g (b) 9.4 g (c) 0.47 g (d) 0.94 g

19. HCl for 5.0 g $MnO_2$:
(a) 2.1 g (b) 8.4 g (c) 16.8 g (d) 4.2 g

20. Average molar mass of argon:
(a) 40.00 (b) 38.0 (c) 39.948 (d) 36.0

*1 (c) · 2 (a) · 3 (d) · 4 (b) · 5 (b) · 6 (d) · 7 (a) · 8 (c) · 9 (d) · 10 (b) · 11 (c) · 12 (a) · 13 (b) · 14 (c) · 15 (d) · 16 (a) · 17 (a) · 18 (d) · 19 (b) · 20 (c)*

## Assertion–Reason (6)

(a) Both true, R explains A. (b) Both true, R does not explain A. (c) A true, R false. (d) A false, R true.

1. A: 0.200 has three significant figures. R: Trailing zeros after a decimal point are significant.
2. A: The molality of a solution is independent of temperature. R: It depends on masses only.
3. A: An element can be broken into simpler substances by chemical methods. R: An element has only one kind of atoms.
4. A: Equal volumes of $H_2$ and $O_2$ at the same T and P have equal numbers of molecules. R: This is Avogadro's law.
5. A: The empirical and molecular formulae of water are the same. R: For water n = 1.
6. A: Mass is conserved in a chemical reaction. R: Atoms are rearranged, not created or destroyed.

*Key: 1 (a) · 2 (a) · 3 (d) · 4 (a) · 5 (a) · 6 (a)*

## Answers to the practice set

**A.** 1. mole 2. $6.022 \times 10^{23}$ 3. carbon-12 4. Proust 5. temperature 6. CH 7. 44.01 g $mol^{-1}$ 8. homogeneous

**B.** 1. F 2. F 3. T 4. F 5. T 6. T 7. F (6.2: the last digit 2 is even) 8. T

**C.** 1. 180.162 u 2. 298.15 K 3. four 4. $1.6 \times 10^{-4}$ 5. 29.25 g 6. Matter is neither created nor destroyed in a chemical reaction 7. sand and salt (oil and water) 8. 58.5 u 9. 0.4 M 10. The reactant used up first, which limits the product

**D.** 1. C 52.14%, H 13.13%, O 34.73% 2. $CH_2Cl$ 3. 36 g 4. 2.79 m 5. 10% 6. Precision: agreement among repeated values; accuracy: agreement with the true value 7. 35.45 u 8. +/−: least decimal places; ×/÷: least significant figures

**E.** 1. $H_2$ limiting; $3.30 \times 10^3$ mol = 56.1 kg $NH_3$ 2. $C_2H_4Cl_2$ 3. Matter → pure substances (elements, compounds) and mixtures (homogeneous, heterogeneous) 4. Indivisible atoms; identical atoms of an element; fixed ratios; reactions rearrange atoms 5. 15.44 M 6. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$; 5 mol $O_2$

**F.** 1. Mole = $6.022 \times 10^{23}$ entities; molar mass in g $mol^{-1}$; mass %, mole fraction, molarity, molality with examples 2. Significant figure rules, arithmetic rules, rounding (even-odd), factor-label method

**G.** 1. It is an average over the isotopes $^{35}Cl$ and $^{37}Cl$ weighted by abundance 2. 1 g $H_2$ (0.5 mol) vs 1 g $O_2$ (0.03125 mol): $H_2$ 3. It depends on masses, not volumes, so it does not change with temperature 4. No: mass is conserved (for nuclear

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reactions mass-energy is conserved instead)

**Case.** (i) 0.5/100 = 0.005 mol (ii) 0.010 mol HCl: 100 mL (iii) 0.005 × 22.7 = 0.114 L (iv) 50 mL has 0.005 mol HCl, needing 0.0025 mol $CaCO_3$: HCl is limiting

## Last-night summary

- Matter: elements, compounds, mixtures; SI units; scientific notation; significant figures and rounding.
- Laws: conservation of mass, definite and multiple proportions, gaseous volumes, Avogadro; Dalton's theory.
- Atomic mass on the $^{12}C$ scale; average atomic mass; molecular and formula mass.
- Mole = $6.022 \times 10^{23}$; n = m/M; empirical vs molecular formula.
- Stoichiometry, limiting reagent; mass %, mole fraction, molarity ($M_1V_1 = M_2V_2$), molality.

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