

CUET · CHEMISTRY · CLASS XI · CODE 306

Some Basic Concepts of Chemistry

CUET unit: Some Basic Concepts of Chemistry

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 **Snapshot**

- Foundation unit of Class XI Chemistry that establishes the **language** of the subject — SI units, scientific notation, significant figures, and dimensional analysis — together with the **quantitative tools** (mole concept, stoichiometry, molarity, molality) used in every subsequent chapter.
- Anchors the five empirical laws of chemical combination — conservation of mass, definite proportions, multiple proportions, Gay-Lussac's law of gaseous volumes, and Avogadro's law — which historically motivated Dalton's atomic theory (1808).
- Defines the mole as the SI unit of amount of substance and fixes the Avogadro number at exactly $6.02214076 \times 10^{23}$ entities per mole (2019 SI redefinition), linking macroscopic mass with microscopic particle count.
- Lays out the empirical → molecular formula pipeline (percentage composition → moles → simplest ratio → multiplier n) that is tested repeatedly in CUET numericals.
- Introduces stoichiometric calculations, the limiting reagent concept, and the four standard concentration measures — mass per cent, mole fraction, molarity, molality — with the $M_1V_1 = M_2V_2$ dilution shortcut.
- Heavy CUET footprint: numerical MCQs on mole–mass–particle interconversion, limiting reagent, molarity, significant figures, density–percentage–to–molarity conversions, and empirical/molecular formula calculations recur every year; expect 8–10 questions per attempt.

 **Detailed Notes****2.1 Core concepts**

Chemistry is the branch of science that studies the preparation, properties, structure and reactions of material substances. It underpins the manufacture of fertilisers, alkali, acids, salts, dyes, polymers, drugs (cisplatin and taxol in cancer therapy, AZT in HIV control), soaps, detergents and metals, and it is central to weather prediction, ozone-layer protection, and pollution control (NCERT §1.1, p. 4).

Nature and states of matter. Matter is anything that has mass and occupies space. At the macroscopic level it occurs in three physical states — solid (definite volume, definite shape, particles close-packed in fixed positions), liquid (definite volume, no

definite shape, particles close but mobile), and gas (no definite volume or shape, particles far apart and moving freely) — and these states are interconvertible by changing temperature or pressure (NCERT §1.2.1, p. 4–5; Fig. 1.1).

Classification of matter. Matter is first classified at the bulk level into pure substances (fixed composition, cannot be separated by physical methods) and mixtures (variable composition, separable by physical methods). Mixtures are homogeneous (sugar in water, air, brass — uniform throughout) or heterogeneous (sugar + salt, grains + pulses + dirt — non-uniform). Pure substances split further into elements (one type of atom — Na, Cu, H₂, O₂) and compounds (two or more elements combined in a fixed ratio — H₂O, CO₂, sugar); a compound's properties differ from those of the elements that make it (NCERT §1.2.2, p. 5–6; Figs. 1.2–1.4).

Properties of matter and SI units. Properties are physical (colour, melting point, boiling point, density — measured without changing chemical identity) or chemical (combustibility, acidity, reactivity — require a chemical change) (NCERT §1.3.1, p. 6). The International System of Units (SI), adopted by the 11th CGPM in 1960 and redefined in 2019, has seven base units — metre (length), kilogram (mass), second (time), ampere (electric current), kelvin (thermodynamic temperature), mole (amount of substance), and candela (luminous intensity) — from which all derived units are built (NCERT §1.3.3, Table 1.1, p. 7). Mass is the amount of matter present and is invariant; weight is the gravitational force on a body and changes with location. Volume has units of (length)³; 1 L = 1000 mL = 1000 cm³ = 1 dm³ (NCERT §1.3.4–1.3.5, p. 9). Density (mass per unit volume) has SI unit kg m⁻³, although chemists routinely use g cm⁻³. The Celsius scale is set by °C = (°F – 32) × 5/9, and the Kelvin scale by K = °C + 273.15; negative readings exist on °C but not on K (NCERT §1.3.6–1.3.7, p. 10; Fig. 1.8).

Uncertainty in measurement. Scientific notation expresses every quantity as $N \times 10^n$ with $1 \leq N < 10$, so that very large (6.022×10^{23}) and very small (1.66×10^{-24} g) numbers fit comfortably on the page and in calculators (NCERT §1.4.1, p. 11). The rules for significant figures are: all non-zero digits are significant; captive zeros between non-zero digits are significant; leading zeros are not; trailing zeros are significant only if a decimal point is present (so 100 has 1 sig fig but 100. has 3 and 100.0 has 4); counted exact numbers have infinite sig figs (NCERT §1.4.2, p. 12). In addition/subtraction, the answer is limited by the fewest decimal places; in multiplication/division, by the fewest significant figures. Precision is the closeness of repeated readings to one another; accuracy is the closeness of a reading to the true value (Table 1.4, p. 13). Dimensional analysis (factor-label method) uses unit-conversion factors equal to unity (e.g. 2.54 cm/1 in) to interconvert units cleanly (NCERT §1.4.3, p. 13–14).

Laws of chemical combination. (i) **Conservation of Mass** (Lavoisier, 1789) — matter is neither created nor destroyed in any physical or chemical change. (ii) **Definite Proportions** (Proust) — a given compound always contains the same elements in the same proportion by mass irrespective of source or method of preparation. (iii) **Multiple Proportions** (Dalton, 1803) — when two elements form more than one compound, the masses of one combining with a fixed mass of the other bear a small whole-number



ratio (H in H_2O vs $\text{H}_2\text{O}_2 \rightarrow \text{O}$ masses $16 : 32 = 1 : 2$). (iv) **Gay-Lussac's Law of Gaseous Volumes** (1808) — gases combine or are produced in simple volume ratios at the same T and P ($\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$ is volumetrically $2 : 1 : 2$). (v) **Avogadro's Law** (1811) — equal volumes of all gases at the same T and P contain equal numbers of molecules (NCERT §1.5, p. 14–16; Fig. 1.9).

Dalton's atomic theory (1808) crystallises these laws: matter is made of indivisible atoms; atoms of an element are identical in mass and properties; compounds form when atoms combine in fixed ratios; chemical reactions only rearrange atoms — never create or destroy them (NCERT §1.6, p. 16).

Atomic, molecular and formula mass. Atomic mass is given relative to ^{12}C , which is assigned exactly 12 u. One unified atomic mass unit (u or amu) = 1.66056×10^{-24} g. The average atomic mass weights isotopic masses by natural abundance (C = 12.011 u, blending ^{12}C , ^{13}C and trace ^{14}C). Molecular mass is the sum of atomic masses ($\text{H}_2\text{O} = 18.02$ u, $\text{CH}_4 = 16.043$ u). For ionic substances without discrete molecules (NaCl), the analogous quantity is called formula mass ($\text{NaCl} = 58.5$ u) (NCERT §1.7, p. 16–18; Fig. 1.10).

The mole. One mole is the SI amount of substance containing exactly $6.02214076 \times 10^{23}$ elementary entities — atoms, molecules, ions, electrons, or specified groups — and this fixed value defines the Avogadro constant N_A . The molar mass in g mol^{-1} is numerically equal to the atomic/molecular/formula mass in u (NCERT §1.8, p. 18; Fig. 1.11). Mole ↔ mass ↔ particle interconversion uses molar mass and N_A .

Percentage composition, empirical and molecular formulas. Mass per cent of an element = (mass of that element in 1 mol / molar mass) $\times 100$. For ethanol $\text{C}_2\text{H}_5\text{OH}$: C = 52.14 %, H = 13.13 %, O = 34.73 %. From percentage composition: convert to moles → divide by smallest mole value → multiply to whole numbers → that ratio is the empirical formula. The molecular formula = (empirical formula) $\times n$, where $n = (\text{molar mass}) / (\text{empirical formula mass})$ (NCERT §1.9–1.9.1, p. 18–20).

Stoichiometry and limiting reagent. A balanced equation gives molar (and hence mass and volume) ratios. For $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$, combustion of 16 g CH_4 yields 36 g H_2O and 44 g CO_2 (Problem 1.3, p. 20–22). The limiting reagent is the reactant present in the least **stoichiometric** quantity — it is consumed first and decides the yield. Problem 1.5 (p. 22): mixing 50.0 kg N_2 with 10.0 kg H_2 to make NH_3 — moles needed of H_2 (5358) exceed moles available (4960), so H_2 is limiting; NH_3 formed = 3.30×10^3 mol = 56.1 kg.

Concentration of solutions. Four common expressions — mass per cent (mass of solute $\times 100$ /mass of solution), mole fraction ($x_a = n_a/(n_a+n_b)$), **molarity** $M = \text{moles of solute} / \text{volume of solution in L}$, and **molality** $m = \text{moles of solute} / \text{mass of solvent in kg}$. The dilution shortcut $M_1V_1 = M_2V_2$ follows from conservation of moles. Molarity changes with temperature because the solution volume expands; molality does not, because the solvent mass is invariant (NCERT §1.10.2, p. 21–24).

2.2 Definitions to memorise

Term	Definition	Page
Matter	Anything that has mass and occupies space	4
Pure substance	Substance with all constituent particles of the same chemical nature and a fixed composition	5
Element	Pure substance made of only one kind of atom (Na, H ₂ , O ₂)	6
Compound	Pure substance formed when two or more elements combine in a fixed mass ratio (H ₂ O, CO ₂)	6
Mass	Amount of matter present in a body — invariant with location	9
Weight	Gravitational force acting on a body — varies with location	9
Density	Mass per unit volume; SI unit kg m ⁻³ , commonly g cm ⁻³	10
Significant figures	Meaningful digits known with certainty plus one estimated (uncertain) digit	12
Precision	Closeness of various measurements of the same quantity to one another	13
Accuracy	Closeness of a measured value to the true (accepted) value	13
Law of Conservation of Mass	In every physical/chemical change there is no net change in mass — Lavoisier, 1789	14
Law of Definite Proportions	A given compound always contains the same elements in the same mass proportion — Proust	15
Law of Multiple Proportions	When two elements form more than one compound, the masses of one combining with a fixed mass of the other are in small whole-number ratios — Dalton, 1803	15
Gay-Lussac's Law	Gases combine in simple volume ratios at the same T and P — 1808	15
Avogadro's Law	Equal volumes of gases at the same T and P contain equal numbers of molecules — 1811	15
Atomic mass unit (u / amu)	One-twelfth of the mass of one ¹² C atom; 1 u = 1.66056 × 10 ⁻²⁴ g	16–17
Average atomic mass	Weighted mean of isotopic masses using natural-abundance fractions	17
Molecular mass	Sum of atomic masses of all atoms in a molecule	17
Formula mass	Analogue of molecular mass for ionic compounds with no discrete molecule (NaCl)	18
Mole	SI unit of amount of substance; contains exactly 6.02214076 × 10 ²³ elementary entities	18

Term	Definition	Page
Molar mass	Mass of one mole of a substance in grams; numerically equal to atomic/molecular/formula mass in u	18
Empirical formula	Simplest whole-number ratio of atoms of different elements in a compound	19
Molecular formula	Actual number of each kind of atom in a molecule of the compound	19
Limiting reagent	Reactant consumed first that thereby limits the amount of product formed	21
Molarity (M)	Number of moles of solute per litre of solution (mol L^{-1})	23
Molality (m)	Number of moles of solute per kilogram of solvent (mol kg^{-1})	24
Mole fraction (x)	Ratio of moles of one component to total moles in the mixture	23

2.3 Diagrams / processes to remember

Make sure you can sketch and label each diagram below. **Fig. 1.1** (p. 5) shows the three states with their characteristic particle arrangement: solids close-packed and ordered, liquids close but disordered, gases far apart and rapidly moving. **Fig. 1.2** (p. 5) is the master classification flowchart — Matter → Pure substance (Elements/Compounds) or Mixture (Homogeneous/Heterogeneous); CUET frequently tests where a given example (air, milk, brass, sodium chloride solution) lands on this tree. **Fig. 1.3** (p. 6) contrasts atoms (Na, Cu) with diatomic molecules (H_2 , N_2 , O_2), and **Fig. 1.4** depicts H_2O and CO_2 molecules — useful for distinguishing elemental versus compound molecules.

Fig. 1.6 (p. 9) is a cube illustrating volume relationships: $1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$. **Fig. 1.7** (p. 10) shows the four volume-measuring devices a CUET candidate must be able to identify by name and by precision class — graduated cylinder, burette, pipette, volumetric flask. **Fig. 1.8** (p. 10) puts $^\circ\text{C}$, $^\circ\text{F}$ and K side by side so that you can compare the freezing point of water ($0 \text{ }^\circ\text{C} / 32 \text{ }^\circ\text{F} / 273.15 \text{ K}$) and human body temperature ($37 \text{ }^\circ\text{C} / 98.6 \text{ }^\circ\text{F} / 310.15 \text{ K}$).

Fig. 1.9 (p. 16) is the central visual of the laws of chemical combination: two volumes of H_2 + one volume of O_2 → two volumes of water vapour, with little boxes showing equal numbers of molecules per equal volumes (Avogadro's bridge from Gay-Lussac). **Fig. 1.10** (p. 17) depicts cubic packing of Na^+ and Cl^- in NaCl — each ion surrounded by six counter-ions — explaining why NaCl has a formula mass rather than a molecular mass. **Fig. 1.11** (p. 18) lays out one mole of various substances (12 g C, 18 g H_2O , 23 g Na, 32 g S, 56 g Fe, 63.5 g Cu, 200 g Hg) side by side to drive home that "one mole" denotes a count of particles, not a fixed mass.

Three workflow processes are worth memorising. The **mole-conversion tree** (p. 18) maps mass ↔ moles ↔ number of particles via molar mass and N_A , and mass ↔ volume via density — this is the spine of every CUET numerical. The **empirical** →

molecular formula process (p. 19–20) takes percentage composition → divide by atomic mass to get moles → divide by the smallest mole to get the empirical ratio → multiply by $n = M_{\text{molar}}/M_{\text{empirical}}$ to get the molecular formula. The **stoichiometry-with-limiting-reagent process** (p. 20–22) is: convert all masses to moles → compare available moles to stoichiometric moles required → identify the limiting reagent → use it to compute product moles → convert back to mass or volume. Practising these three procedural pipelines once each will solve >70 % of CUET kech101 numericals.

2.4 Common confusions / NTA trap points

- **Mass vs weight** — students treat them as synonyms; NCERT explicitly states (p. 9) that mass is constant while weight varies with gravitational acceleration. Distractor: "weight is the amount of matter".
- **Significant figures with terminal zeros** — 100 has only 1 sig fig, 100. has 3, 100.0 has 4 (p. 12). NTA loves this trap; the decimal point is decisive.
- **Definite vs Multiple Proportions** — "Definite" is about ONE compound (always the same composition); "Multiple" is about TWO or more compounds of the same two elements (whole-number mass ratios). Don't confuse the two laws.
- **Gay-Lussac vs Avogadro** — Gay-Lussac talks about combining **volumes**; Avogadro extends it to equal **molecules** per equal volume. They are different statements, and CUET MCQs deliberately swap the names.
- **Molarity vs molality** — molarity depends on solution volume and so varies with temperature; molality uses solvent mass and so is temperature-independent (p. 24).
- **Limiting-reagent shortcut** — compare moles **available** to moles **required** by stoichiometry; the reactant that yields the **fewer** moles of product is limiting. Smaller mass alone is NOT the criterion (Problem 1.5).
- **amu definition** — exactly $1/12$ the mass of one ^{12}C atom (p. 16); not "the mass of a hydrogen atom" (that is the older 1803 system).
- **Empirical vs molecular formula** — glucose has empirical formula CH_2O but molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$. CUET asks you to find $n = \text{molar mass} / \text{empirical formula mass}$; don't forget to multiply through.
- **Mole fraction has no units** — it is a ratio of moles, so its sum over all components must equal 1. Trap: choosing % composition as a "mole fraction".
- **Avogadro number is now exact** — since 2019 it is fixed at $6.02214076 \times 10^{23} \text{ mol}^{-1}$ by definition, not measured (p. 18). NTA may phrase this as "approximate" in a distractor.
- **Counting particles in diatomic molecules** — 1 mole of Cl_2 contains 6.022×10^{23} molecules but $2 \times 6.022 \times 10^{23}$ Cl atoms. Watch the wording.
- **Density-percentage-to-molarity** — $M = (10 \times \text{density} \times \% \text{ w/w}) / \text{molar mass}$ is a derived shortcut for problems where concentrated solutions are described by weight percent and density (Exercise 1.38, p. 29).

2.5 Key reactions & formulas

Reaction / Formula	Conditions / Notes	NCERT page
$1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$	Volume conversions	9
Density = mass / volume	SI kg m^{-3} ; chemistry g cm^{-3}	10
$K = ^\circ\text{C} + 273.15$	Kelvin from Celsius	10
$^\circ\text{F} = (9/5)^\circ\text{C} + 32$	Fahrenheit from Celsius	10
$1 \text{ u} = 1.66056 \times 10^{-24} \text{ g}$	Atomic mass unit, ^{12}C reference	17
$N_A = 6.02214076 \times 10^{23} \text{ mol}^{-1}$	Avogadro constant (exact since 2019)	18
$n \text{ (moles)} = \text{mass (g)} / \text{molar mass (g mol}^{-1}\text{)}$	Mass-to-mole	18
Number of particles = $n \times N_A$	Mole-to-particle	18
Mass per cent = (mass of element / molar mass) $\times 100$	Element composition in a compound	18
$n \text{ (multiplier)} = M_{\text{molar}} / M_{\text{empirical formula}}$	Empirical \rightarrow molecular formula	19
$2 \text{ H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{ H}_2\text{O}(\text{l})$	Combustion of hydrogen — illustrates Gay-Lussac/Avogadro	15–16
$\text{CH}_4(\text{g}) + 2 \text{ O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{ H}_2\text{O}(\text{l})$	Stoichiometry worked example (Problem 1.3)	20
$\text{N}_2(\text{g}) + 3 \text{ H}_2(\text{g}) \rightarrow 2 \text{ NH}_3(\text{g})$	Limiting-reagent worked example (Problem 1.5)	22
$\text{CaCO}_3(\text{s}) + 2 \text{ HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	Acid-carbonate stoichiometry (Exercise 1.35)	28
Molarity (M) = $n_{\text{solute}} / V_{\text{solution}} \text{ (L)}$	Mol L^{-1} ; temperature-dependent	23
Molality (m) = $n_{\text{solute}} / \text{mass}_{\text{solvent}} \text{ (kg)}$	Mol kg^{-1} ; temperature-independent	24
Mole fraction $x_A = n_A / (n_A + n_B)$	$\sum x_i = 1$; dimensionless	23
$M_1V_1 = M_2V_2$	Dilution; moles conserved	24
$M = (10 \times d \times \% \text{ w/w}) / M_{\text{molar}}$	Density-percentage shortcut for molarity (g cm^{-3} , %)	29
Average atomic mass = $\sum (f_i \times m_i)$	Weighted by natural abundance f_i	17



Practice MCQs

Q1. How many significant figures are present in the measurement 0.00500400?

- A. 4
- B. 5
- C. 6
- D. 8

Q2. 50.0 kg of $\text{N}_2(\text{g})$ is mixed with 10.0 kg of $\text{H}_2(\text{g})$ for $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$. The limiting reagent and the mass of NH_3 produced are, respectively:

- A. N_2 ; 60.7 kg
- B. H_2 ; 56.1 kg
- C. N_2 ; 56.1 kg
- D. H_2 ; 60.7 kg

Q3. The molarity of a solution prepared by dissolving 4.0 g of NaOH in enough water to give 250 mL of solution is: ($M_{\text{NaOH}} = 40 \text{ g mol}^{-1}$)

- A. 0.1 M
- B. 0.2 M
- C. 0.4 M
- D. 1.0 M

 **12 more MCQs + answer key**

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PYQ Alignment

This chapter is among the most heavily tested in CUET Chemistry, contributing roughly **8–10 MCQs every year** across CUET 2023–25. The favourite themes are: significant-figure rules; mole–mass–particle interconversion (especially the "largest number of atoms in 1 g of ..." style); limiting reagent numericals; molarity/molality/dilution calculations

using $M_1V_1 = M_2V_2$; identifying the law of chemical combination from given data; empirical-to-molecular formula problems; and matching SI base units to base quantities.

For chapter-wise PYQ practice and consolidated mock tests on these themes, see [/pyq/chemistry](#).

CUET 2024 — Actual PYQs from this chapter

Q.2 (CUET 2024) Arrange the following in decreasing order of number of molecules contained in: (A) 16 g of O_2 (B) 16 g of CO_2 (C) 16 g of CO (D) 16 g of H_2

- A) (A), (B), (C), (D) B) (D), (C), (A), (B) C) (B), (A), (D), (C) D) (C), (B), (D), (A)

Tests: Mole concept & stoichiometry **Answer:** Not in extracted key

Q.32 (CUET 2024) Number of moles of hydrogen oxidised is:

- A) 0.33 mol B) 33.3 mol C) 3 mol D) 1.33 mol

Tests: Mole concept & stoichiometry **Answer:** Not in extracted key

