

CHEMISTRY

Chapter 1 — Some Basic Concepts of Chemistry

Class 11 | NEET · IIT-JEE · Board Examination Prep

COMPLETE QUESTION BANK

30 MCQs | 20 Fill in the Blanks | 10 Matching

30 FAQs | 20 Short Q&A | 10 Long Answers

Total: 120 Questions with Full Answers & Hints

All answers are provided immediately after each question

30 MCQs	20 Fill Blanks	10 Matching	30 FAQs	20 Short Q&A	10 Long Ans
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SECTION 1: MULTIPLE CHOICE QUESTIONS (MCQs)

30 Questions | 1 mark each | 4 options per question | Correct answer + Hint provided

Q1. Which branch of chemistry studies carbon-based compounds?

- (A) Inorganic chemistry
- (B) Organic chemistry
- (C) Physical chemistry
- (D) Nuclear chemistry

✓ **Answer: (B) Organic chemistry**

■ *Hint: Carbon is the backbone of life — organic chemistry covers millions of carbon compounds.*

Q2. Which of the following is a CHEMICAL change?

- (A) Melting of ice
- (B) Dissolution of sugar in water
- (C) Rusting of iron
- (D) Cutting of paper

✓ **Answer: (C) Rusting of iron**

■ *Hint: Rust = Fe_2O_3 , a NEW substance. Irreversible — confirms chemical change.*

Q3. The SI unit of amount of substance is:

- (A) Gram
- (B) Kilogram
- (C) Mole
- (D) Litre

✓ **Answer: (C) Mole**

■ *Hint: 1 mole = 6.022×10^{23} particles. Mole is the chemist's counting unit.*

Q4. Avogadro's number (N_A) equals:

- (A) 6.022×10^{22}
- (B) 6.022×10^{23}
- (C) 6.022×10^{24}
- (D) 6.022×10^{20}

✓ **Answer: (B) 6.022×10^{23}**

■ *Hint: This is the number of particles in exactly 1 mole of any substance.*

Q5. Molar mass of water (H₂O) is:

- (A) 16 g/mol
- (B) 18 g/mol
- (C) 20 g/mol
- (D) 17 g/mol

✓ **Answer: (B) 18 g/mol**

■ *Hint: H = 1×2 = 2; O = 16. Total = 18 g/mol. The most important molar mass to remember!*

Q6. 0°C equals how many Kelvin?

- (A) 0 K
- (B) 100 K
- (C) 273 K
- (D) 373 K

✓ **Answer: (C) 273 K**

■ *Hint: Formula: K = °C + 273. Always add 273 when converting Celsius to Kelvin.*

Q7. Which law states "mass is neither created nor destroyed in a chemical reaction"?

- (A) Law of Definite Proportions
- (B) Law of Conservation of Mass
- (C) Law of Multiple Proportions
- (D) Gay-Lussac's Law

✓ **Answer: (B) Law of Conservation of Mass**

■ *Hint: Lavoisier's law: mass before reaction = mass after reaction.*

Q8. Water always contains H and O in mass ratio 1:8. This is an example of:

- (A) Law of Conservation of Mass
- (B) Law of Multiple Proportions
- (C) Law of Definite Proportions
- (D) Gay-Lussac's Law

✓ **Answer: (C) Law of Definite Proportions**

■ *Hint: Proust's law: a compound always has the same element ratio regardless of source.*

Q9. Number of atoms in 12 g of carbon-12 is:

- (A) 6.022×10^{23}
- (B) 3.011×10^{23}
- (C) 12.044×10^{23}
- (D) 1.204×10^{23}

✓ **Answer: (A) 6.022×10^{23}**

■ *Hint: 12 g of C-12 = exactly 1 mole = N_A atoms. This is by definition!*

Q10. Which compound has the HIGHEST molar mass?

- (A) NaCl (58.5)
- (B) H_2SO_4 (98)
- (C) Glucose $C_6H_{12}O_6$ (180)
- (D) CO_2 (44)

✓ **Answer: (C) Glucose $C_6H_{12}O_6 = 180$ g/mol**

■ *Hint: $6(12) + 12(1) + 6(16) = 72 + 12 + 96 = 180$ g/mol.*

Q11. Number of significant figures in 0.00340 is:

- (A) 5
- (B) 3
- (C) 6
- (D) 2

✓ **Answer: (B) 3**

■ *Hint: Leading zeros are NOT significant. Trailing zero after decimal IS. So: 3, 4, 0 → 3 sig figs.*

Q12. Number of molecules in 9 g of water:

- (A) 6.022×10^{23}
- (B) 3.011×10^{23}
- (C) 18.066×10^{23}
- (D) 1.505×10^{23}

✓ **Answer: (B) 3.011×10^{23}**

■ *Hint: $n = 9/18 = 0.5$ mol. Molecules = $0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$.*

Q13. Which is a HOMOGENEOUS mixture?

- (A) Sand and water
- (B) Salt water (saline)
- (C) Iron filings and sulfur
- (D) Sand and gravel

✓ **Answer: (B) Salt water (saline)**

■ *Hint: Salt dissolves completely — uniform composition throughout. That's homogeneous.*

Q14. CO and CO₂ are examples illustrating which law?

- (A) Law of Definite Proportions
- (B) Law of Conservation of Mass
- (C) Law of Multiple Proportions
- (D) Law of Reciprocal Proportions

✓ **Answer: (C) Law of Multiple Proportions**

■ *Hint: Same two elements (C, O), two compounds. O ratio per fixed C = 16:32 = 1:2 (simple whole number).*

Q15. 1 mole of any ideal gas at STP occupies:

- (A) 11.2 L
- (B) 22.4 L
- (C) 44.8 L
- (D) 2.24 L

✓ **Answer: (B) 22.4 L**

■ *Hint: Molar volume at STP (0°C, 1 atm) = 22.4 L/mol. MUST memorise for JEE/NEET.*

Q16. Which postulate of Dalton's theory was later found INCORRECT?

- (A) Atoms combine in simple whole number ratios
- (B) Atoms are indivisible
- (C) Atoms rearrange in chemical reactions
- (D) Atoms of same element have same chemical properties

✓ **Answer: (B) Atoms are indivisible**

■ *Hint: Sub-atomic particles (e^- , p^+ , n) were later discovered inside atoms.*

Q17. Number of moles in 44 g of CO₂:

- (A) 0.5 mol
- (B) 1 mol
- (C) 2 mol
- (D) 1.5 mol

✓ **Answer: (B) 1 mol**

■ *Hint: Molar mass CO₂ = 12+32 = 44 g/mol. $n = 44/44 = 1$ mol.*

Q18. Scientific notation of 0.000456 is:

- (A) 4.56×10^{-3}
- (B) 4.56×10^{-4}
- (C) 4.56×10^{-5}
- (D) 456×10^{-6}

✓ **Answer: (B) 4.56×10^{-4}**

■ *Hint: Move decimal 4 places right → exponent = -4. N must be between 1 and 10.*

Q19. The limiting reagent in a reaction is the reactant that:

- (A) Has the largest molar mass
- (B) Is consumed completely first and limits product amount
- (C) Is always taken in excess
- (D) Produces the maximum amount of product

✓ **Answer: (B) Is consumed completely first and limits product amount**

■ *Hint: Like "chai patti" — when it runs out, no more tea, regardless of milk/water left!*

Q20. Isotopes are atoms of the same element differing in:

- (A) Number of protons
- (B) Atomic number
- (C) Number of neutrons (and thus mass number)
- (D) Chemical properties

✓ **Answer: (C) Number of neutrons (and thus mass number)**

■ *Hint: Example: C-12 and C-14. Same element (same protons), different mass (different neutrons).*

Q21. Which of the following is a PURE substance?

- (A) Air
- (B) Sea water
- (C) Diamond
- (D) Stainless steel

✓ **Answer: (C) Diamond**

■ *Hint: Diamond = pure carbon only. Air, sea water, and steel are all mixtures.*

Q22. Mass of one atom of hydrogen is approximately:

- (A) 1.66×10^{-24} g
- (B) 1.66×10^{-23} g
- (C) 1.66×10^{-22} g
- (D) 1.66×10^{-20} g

✓ **Answer: (A) 1.66×10^{-24} g**

■ *Hint: $1 \text{ u} = 1.66 \times 10^{-24}$ g. H has atomic mass $\approx 1 \text{ u}$, so mass $\approx 1.66 \times 10^{-24}$ g.*

Q23. Which is a PHYSICAL change?

- (A) Burning of coal
- (B) Rusting of iron
- (C) Melting of wax
- (D) Curdling of milk

✓ **Answer: (C) Melting of wax**

■ *Hint: Wax changes state (solid \rightarrow liquid) but remains wax chemically. Reversible on cooling.*

Q24. Gay-Lussac's Law of Gaseous Volumes states that:

- (A) Masses of reactants are in whole number ratios
- (B) Volumes of gaseous reactants/products are in simple whole number ratios
- (C) Temperature and pressure are proportional
- (D) Pressure and volume are inversely proportional

✓ **Answer: (B) Volumes of gaseous reactants/products are in simple whole number ratios**

■ *Hint: Example: $2 \text{ vol } H_2 + 1 \text{ vol } O_2 \rightarrow 2 \text{ vol } H_2O$. Ratio = 2:1:2.*

Q25. Molar mass of H_2SO_4 is:

- (A) 96 g/mol
- (B) 98 g/mol
- (C) 97 g/mol
- (D) 100 g/mol

✓ **Answer: (B) 98 g/mol**

■ *Hint: $2(1) + 32 + 4(16) = 2+32+64 = 98$ g/mol. Important acid for NEET/JEE problems.*

Q26. Percentage yield formula is:

- (A) $(\text{Theoretical}/\text{Actual}) \times 100$
- (B) $(\text{Actual}/\text{Theoretical}) \times 100$
- (C) $(\text{Actual} - \text{Theoretical}) \times 100$
- (D) $\text{Actual} \times \text{Theoretical}$

✓ **Answer: (B) $(\text{Actual}/\text{Theoretical}) \times 100$**

■ *Hint: Always $\leq 100\%$. Actual = what you get in lab; Theoretical = max possible from stoichiometry.*

Q27. Which ancient Indian scholar proposed the concept of "parmanu" ~2600 years ago?

- (A) Aryabhata
- (B) Brahmagupta
- (C) Acharya Kanad
- (D) Chanakya

✓ **Answer: (C) Acharya Kanad**

■ *Hint: Kanad called indivisible particles "parmanu" (■■■■■■■). Dalton proposed similar ideas 2400 years later!*

Q28. Formula mass of NaCl is:

- (A) 58 u
- (B) 58.5 u
- (C) 59 u
- (D) 57.5 u

✓ **Answer: (B) 58.5 u**

■ *Hint: Na = 23, Cl = 35.5. Total = 58.5 u. Ionic compounds use "formula mass", not molecular mass.*

Q29. Which of the following has exactly 3 significant figures?

- (A) 0.003
- (B) 103.0
- (C) 1030
- (D) 3.00

✓ **Answer: (D) 3.00**

■ *Hint: 3.00 → digits 3, 0, 0. Trailing zeros after decimal ARE significant. So = 3 sig figs.*

Q30. At absolute zero (0 K), what is theoretically true?

- (A) All molecules move at maximum speed
- (B) All molecular motion ceases
- (C) Gas expands to infinite volume
- (D) All substances become gaseous

✓ **Answer: (B) All molecular motion ceases**

■ *Hint: 0 K = -273°C. Theoretically, no kinetic energy remains. Cannot go lower than 0 K.*

SECTION 2: FILL IN THE BLANKS

20 Questions | Complete each statement with the correct term or value

1. 1 mole of any substance contains _____ particles.

✓ Answer: 6.022×10^{23} (Avogadro's Number, N_A)

2. The SI unit of temperature used in scientific calculations is _____.

✓ Answer: Kelvin (K)

3. Celsius to Kelvin conversion formula: $K =$ _____.

✓ Answer: $^{\circ}\text{C} + 273$

4. Molar mass of CO_2 is _____ g/mol.

✓ Answer: 44 g/mol (C=12, O=16×2=32, Total=44)

5. Law of _____ states that matter can neither be created nor destroyed.

✓ Answer: Conservation of Mass (given by Lavoisier)

6. The reference standard for atomic mass scale is _____ with mass = 12 u.

✓ Answer: Carbon-12 (^{12}C)

7. 1 mole of a gas at STP occupies _____ litres.

✓ Answer: 22.4 litres (Molar Volume at STP)

8. Number of significant figures in 205.0 is _____.

✓ Answer: 4 (2, 0, 5, and the trailing decimal zero are all significant)

9. The reactant that is consumed completely first in a reaction is called _____.

✓ Answer: Limiting Reagent

10. Atoms of the same element having the same atomic number but different mass numbers are _____.

✓ Answer: Isotopes (e.g., C-12 and C-14)

11. Molar mass of H_2SO_4 is _____ g/mol.

✓ Answer: 98 g/mol (2+32+64=98)

12. % Yield = (Actual yield ÷ Theoretical yield) × _____.

✓ Answer: 100

13. Scientific notation of 6,022,000 is _____.

✓ Answer: 6.022×10^6

14. Water (H_2O) always contains H and O in mass ratio _____.

✓ Answer: 1 : 8 (by mass)

15. Formula mass of NaCl = _____ u.

✓ Answer: 58.5 u (Na=23 + Cl=35.5)

16. 0 Kelvin is called _____ and equals _____ °C.

✓ Answer: Absolute Zero ; -273°C

17. A mixture having uniform composition throughout is called a _____ mixture.

✓ Answer: Homogeneous

18. Number of moles (n) = Given mass (w) ÷ _____.

✓ Answer: Molar Mass (M) → $n = w/M$

19. Dalton's Atomic Theory was proposed in the year _____.

✓ Answer: 1808

20. The ancient Indian scholar who proposed "parmanu" concept was _____.

✓ Answer: Acharya Kanad (~600 BCE)

SECTION 3: MATCHING TYPE QUESTIONS

10 Sets | Match Column A with Column B | Answers provided below each set

Set 1 — Mole Concept

Column A (Term)	Column B (Value / Description)
A. Mole	P. 6.022×10^{23} particles
B. Avogadro's Number	Q. g/mol
C. Molar Mass unit	R. SI unit for amount of substance
D. 1 u (atomic mass unit)	S. 1.66×10^{-24} g
Answer:	A → R, B → P, C → Q, D → S

Set 2 — Laws & Scientists

Column A (Law)	Column B (Scientist)
A. Conservation of Mass	P. John Dalton
B. Definite Proportions	Q. Gay-Lussac
C. Multiple Proportions	R. Antoine Lavoisier
D. Gaseous Volumes	S. Joseph Proust
Answer:	A → R, B → S, C → P, D → Q

Set 3 — Compounds & Molar Masses

Column A (Compound)	Column B (Molar Mass)
A. H_2O	P. 44 g/mol
B. CO_2	Q. 180 g/mol
C. NaCl	R. 18 g/mol
D. Glucose $C_6H_{12}O_6$	S. 58.5 g/mol
Answer:	A → R, B → P, C → S, D → Q

Set 4 — Types of Change

Column A (Change)	Column B (Type)
A. Melting of ice	P. Chemical change
B. Rusting of iron	Q. Physical change
C. Burning of coal	R. Physical change

D. Cutting of paper	S. Chemical change
Answer:	A → Q/R, B → P, C → S, D → Q/R

Set 5 — SI Units

Column A (Symbol)	Column B (Physical Quantity)
A. kg	P. Temperature
B. mol	Q. Mass
C. K	R. Length
D. m	S. Amount of substance
Answer:	A → Q, B → S, C → P, D → R

Set 6 — States of Matter

Column A (State)	Column B (Property)
A. Solid	P. No fixed shape or volume
B. Liquid	Q. Fixed shape and fixed volume
C. Gas	R. Fixed volume, variable shape
D. Plasma	S. Ionised gas state
Answer:	A → Q, B → R, C → P, D → S

Set 7 — Classification of Matter

Column A (Term)	Column B (Example)
A. Element	P. NaCl
B. Compound	Q. Air ($N_2 + O_2$)
C. Heterogeneous mixture	R. Gold (Au)
D. Homogeneous mixture	S. Sand in water
Answer:	A → R, B → P, C → S, D → Q

Set 8 — Mole Formulas

Column A (Quantity)	Column B (Formula)
A. Moles (n)	P. $n \times N_A$
B. Number of particles	Q. w / M
C. Volume of gas at STP	R. $n \times 22.4 \text{ L}$

D. Mass from moles	S. $n \times M$
Answer:	A → Q, B → P, C → R, D → S

Set 9 — Key Concepts

Column A (Concept)	Column B (Description)
A. Limiting reagent	P. Same element, different mass number
B. Isotopes	Q. Reactant consumed first
C. Significant figures	R. Precision of measurement
D. Law of Multiple Proportions	S. Whole number ratio of masses
Answer:	A → Q, B → P, C → R, D → S

Set 10 — Temperature Values

Column A (Temperature)	Column B (Significance)
A. 0 K	P. Boiling point of water
B. 273 K	Q. Absolute zero (-273°C)
C. 373 K	R. Standard lab temperature
D. 298 K	S. Melting point of ice (0°C)
Answer:	A → Q, B → S, C → P, D → R

SECTION 4: FREQUENTLY ASKED QUESTIONS (FAQs)

30 Questions | Conceptual clarity | 2–3 marks each | NEET/JEE/Board

Q1. What is chemistry? Give its official definition.

Chemistry is the branch of science that studies the **composition, structure, properties, and reactions** of matter. In simple words: it is the science of substances — what they are made of, how they behave, and how they change. From making tea to digesting food, chemistry is everywhere!

Q2. What is matter? How is it defined in chemistry?

Matter is **anything that has mass and occupies space**. Examples: water, iron, air, food, even perfume (molecules in air). Note: energy (light, heat) is NOT matter since it has no mass.

Q3. Distinguish between physical and chemical changes with two examples each.

Physical change: Only form/state changes; substance identity unchanged; usually reversible.

Examples: (1) Ice melting → water. (2) Cutting paper into pieces.

Chemical change: New substance forms; usually irreversible; composition changes.

Examples: (1) Iron rusting → Fe_2O_3 . (2) Burning coal → CO_2 + ash.

Q4. What are the indicators (signs) of a chemical change?

Signs that a chemical change has occurred: **(1)** New substance formed with different properties. **(2)** Colour change. **(3)** Gas production (bubbles or odour). **(4)** Heat or light released/absorbed. **(5)** Change in smell. **(6)** Usually irreversible. If you see any of these → likely a chemical change!

Q5. Differentiate between an element and a compound.

Element: Pure substance made of only ONE type of atom. Cannot be broken down chemically. 118 elements exist in the Periodic Table. Examples: Gold (Au), Oxygen (O_2), Carbon (C).

Compound: Two or more elements chemically combined in a FIXED ratio. New properties arise. Can be split only by chemical methods. Examples: H_2O , NaCl, CO_2 .

Q6. How does a mixture differ from a compound?

Mixture: Components mixed physically in any ratio; each retains original properties; separated by physical methods (filtration, evaporation).

Compound: Elements combined chemically in fixed ratio; completely new properties; separated only by chemical methods (electrolysis). Example: Salt water = mixture (separated by evaporation). Water = compound (separated by electrolysis).

Q7. What is the SI system? Name the 7 base units.

SI (Système International d'Unités) is the internationally agreed system of measurement. It ensures all scientists use a common language.

7 Base Units: Mass → kg, Length → m, Time → s, Temperature → K, Amount of substance → mol, Electric current → A, Luminous intensity → cd.

Q8. Distinguish between mass and weight.

Mass: Amount of matter in a body. Measured in kg. CONSTANT everywhere (same on Earth, Moon, Mars).

Weight: Force due to gravity on the body. $W = m \times g$. VARIES with location (on Moon, weight = 1/6 of Earth weight, but mass stays same).

In chemistry, we always work with MASS, not weight.

Q9. Why is Kelvin scale preferred over Celsius for scientific calculations?

Kelvin starts at **Absolute Zero (0 K = -273°C)** — the lowest possible temperature. It has no negative values, making it ideal for gas law equations (Boyle's Law, Charles' Law). Using Celsius in gas laws would give wrong answers because negative temperatures would create mathematical errors.

Q10. What is scientific notation? Why is it used? Give two examples.

Scientific notation expresses numbers as $N \times 10^n$ where $1 \leq N < 10$.

Why used: Chemistry involves extremely large and small numbers (e.g., atom masses are $\sim 10^{-24}$ g; Avogadro's number = 6×10^{23}). Writing/calculating these in full is impractical.

Examples: $0.000166 = 1.66 \times 10^{-4}$; $6022000 = 6.022 \times 10^6$.

Q11. What are significant figures? Why do they matter?

Significant figures (sig figs) indicate the **precision of a measurement** — how many digits are reliably known. "5 cm" (1 sig fig) is less precise than "5.00 cm" (3 sig figs). They prevent overstating accuracy. In calculations, the answer cannot be more precise than the least precise measurement used.

Q12. State and explain the Law of Conservation of Mass.

"In any chemical reaction, the total mass of reactants equals the total mass of products. Matter is neither created nor destroyed — only rearranged." — Antoine Lavoisier, 1789.

Example: When iron burns in oxygen: $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$. Mass of Fe + mass of O_2 used = mass of Fe_2O_3 formed.

Q13. State and explain the Law of Definite Proportions.

"A pure compound always contains the same elements combined in the same fixed ratio by mass, regardless of its source or method of preparation." — Joseph Proust, 1799.

Example: H_2O always has H:O = 1:8 by mass. Whether from a river, raincloud, or lab — this ratio never changes.

Q14. State the Law of Multiple Proportions and give an example.

"When two elements form more than one compound, the masses of one element that combine with a fixed mass of the other are in simple whole number ratios." — John Dalton.

Example: CO: 12g C + 16g O. CO_2 : 12g C + 32g O. Oxygen ratio = 16:32 = 1:2 ✓ (simple whole number ratio).

Q15. What is Gay-Lussac's Law of Gaseous Volumes?

"At constant temperature and pressure, volumes of gaseous reactants and products are in simple whole number ratios."

Example: $2 \text{ vol H}_2 + 1 \text{ vol O}_2 \rightarrow 2 \text{ vol H}_2\text{O vapour}$. Ratio = 2:1:2.

Significance: This law led directly to Avogadro's hypothesis about equal volumes of gases containing equal number of molecules.

Q16. State Dalton's Atomic Theory — list all five postulates.

- (1) All matter consists of tiny indivisible particles called ATOMS.
- (2) All atoms of a given element are identical in mass and chemical properties.
- (3) Atoms of different elements differ in mass and properties.
- (4) Atoms combine in simple whole number ratios to form compounds.
- (5) In reactions, atoms are only rearranged — never created or destroyed.

Q17. What are the three main limitations of Dalton's Atomic Theory?

- (1) **Atoms ARE divisible:** Electrons, protons, neutrons were later discovered inside atoms.
- (2) **Isotopes exist:** Atoms of the same element can have different masses (e.g., C-12 and C-14 — same element, different neutron count).
- (3) **Non-stoichiometric compounds:** Some compounds do NOT have simple whole number ratios.

Q18. What is atomic mass? What is the standard reference?

Atomic mass is the **relative mass of an atom** compared to 1/12th the mass of a carbon-12 atom.
Unit: u (unified mass unit) or amu.

Carbon-12 is the reference standard (atomic mass = exactly 12 u).

$1 \text{ u} = 1.66 \times 10^{-24} \text{ g}$ (extremely small — hence we use relative values).

Examples: H=1u, C=12u, O=16u, Na=23u, Cl=35.5u, Fe=56u.

Q19. Distinguish between molecular mass and formula mass.

Molecular mass: Used for covalent compounds that exist as discrete molecules. Sum of atomic masses in one molecule.

Example: $\text{H}_2\text{O} = 2(1)+16 = 18 \text{ u}$.

Formula mass: Used for ionic compounds (no discrete molecules exist).

Example: $\text{NaCl} = 23+35.5 = 58.5 \text{ u}$. Same calculation — just different name for ionic compounds.

Q20. What is a mole? Why is it a useful unit in chemistry?

A **mole** is the SI unit for amount of substance. **1 mole = 6.022×10^{23} particles** (Avogadro's Number).

It is useful because atoms are too tiny to count individually. Just as we count eggs in "dozens" (12), chemists count atoms/molecules in "moles" (6.022×10^{23}). It bridges the atomic scale with lab-measurable quantities.

Q21. What is molar mass? How does it relate to molecular mass?

Molar mass is the mass of 1 mole of a substance, expressed in g/mol. Numerically, it equals the molecular/atomic mass in "u".

Example: H_2O has molecular mass = 18 u \rightarrow molar mass = 18 g/mol.

This means: 18 g of water contains exactly 6.022×10^{23} molecules of water.

Q22. What is molar volume? What is its value at STP?

Molar volume = volume occupied by 1 mole of an ideal gas.

At **STP (0°C, 1 atm)**: Molar volume = **22.4 L/mol** for any ideal gas.

This means 2 g of H_2 , 32 g of O_2 , or 44 g of CO_2 all occupy the same volume: 22.4 L at STP!

Q23. Write the formula to calculate moles. Give one worked example.

Formula: $n = w / M$

where n = moles, w = given mass (g), M = molar mass (g/mol).

Example: Find moles in 36 g of water (H_2O , $M=18$ g/mol).

$n = 36/18 = 2$ moles of water.

Particles = $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$ molecules.

Q24. What is stoichiometry? What does a balanced equation tell us?

Stoichiometry (Greek: stoicheion=element, metron=measure) is the quantitative study of reactants and products in chemical reactions.

A balanced equation tells us: (1) Which substances react and form. (2) Mole ratios of all species. (3) Mass ratios (using molar masses). (4) Volume ratios for gases at STP.

Example: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ tells us 1 mol N_2 needs 3 mol H_2 to give 2 mol NH_3 .

Q25. What is the limiting reagent? How do you identify it?

The **limiting reagent** is the reactant completely consumed first, determining the maximum product formed. The other reactant is in "excess" and some remains unused.

How to identify:

Step 1: Convert all given masses to moles.

Step 2: Divide each by its stoichiometric coefficient from the balanced equation.

Step 3: The reactant giving the SMALLEST value is the limiting reagent.

Q26. What is percentage yield? Why is actual yield less than theoretical?

% Yield = (Actual Yield / Theoretical Yield) \times 100

Theoretical yield = maximum possible product calculated from limiting reagent.

Actual yield = product actually obtained in lab.

Why less? (1) Side reactions consume reactants. (2) Incomplete reaction. (3) Product lost during separation/purification. (4) Measurement errors. Always $\leq 100\%$.

Q27. What was Acharya Kanad's contribution to atomic theory?

Acharya Kanad (~600 BCE) proposed that all matter is made of tiny, **indivisible particles called "Parmanu"** (परमाणु) — meaning "that which cannot be cut further". This is remarkably similar to Dalton's atomic theory proposed ~2400 years later! While Kanad's idea was philosophical, Dalton's was based on experimental evidence.

Q28. State the rules for counting significant figures (all 5 rules).

Rule 1: All non-zero digits are significant. (234 → 3 sig figs)

Rule 2: Zeros BETWEEN non-zero digits are significant. (205 → 3; 10.07 → 4)

Rule 3: LEADING zeros are NOT significant. (0.0045 → 2 sig figs)

Rule 4: TRAILING zeros after decimal ARE significant. (2.500 → 4 sig figs)

Rule 5: Trailing zeros in whole numbers are ambiguous — use scientific notation. ($2500 = 2.5 \times 10^3$ for 2 sig figs)

Q29. What is absolute zero? Why can't temperature go below it?

Absolute zero = 0 K = -273°C. It is the theoretically lowest possible temperature, at which all molecular motion (kinetic energy) completely ceases. Temperature is a measure of molecular motion — since motion cannot be less than zero, 0 K is the absolute minimum. In practice, scientists have reached temperatures extremely close to 0 K but never exactly 0 K.

Q30. Why is chemistry often called the "central science"?

Chemistry connects and overlaps with virtually every other science:

- **Physics:** Nuclear chemistry, quantum mechanics
- **Biology:** Biochemistry, metabolism, DNA
- **Geology:** Mineral formation, geochemistry
- **Medicine:** Drug design, diagnostics
- **Environment:** Pollution, climate change

Every manufactured product, every biological process, every material involves chemistry — hence it is the "central science".

SECTION 5: SHORT QUESTION & ANSWERS

20 Questions | 2–3 marks each | Includes numerical problems | NEET/JEE style

Q1. Calculate the molar mass of glucose ($C_6H_{12}O_6$).

$$C: 6 \times 12 = 72$$

$$H: 12 \times 1 = 12$$

$$O: 6 \times 16 = 96$$

$$\text{Molar mass} = 72 + 12 + 96 = 180 \text{ g/mol}$$

Q2. How many molecules are present in 36 g of water ($M = 18 \text{ g/mol}$)?

$$n = 36/18 = 2 \text{ mol}$$

$$\text{Molecules} = 2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24} \text{ molecules}$$

Q3. Convert: (a) 25°C to Kelvin (b) 100°C to Kelvin (c) 0 K to $^\circ\text{C}$

$$(a) K = 25 + 273 = 298 \text{ K (standard lab/room temperature)}$$

$$(b) K = 100 + 273 = 373 \text{ K (boiling point of water)}$$

$$(c) ^\circ\text{C} = 0 - 273 = -273^\circ\text{C (absolute zero)}$$

Q4. How many atoms are in 24 g of carbon? (Atomic mass of C = 12)

$$n = 24/12 = 2 \text{ mol}$$

$$\text{Atoms} = 2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24} \text{ atoms}$$

Q5. Express in scientific notation: (a) 3,500,000 (b) 0.000789 (c) speed of light = 300,000,000 m/s

$$(a) 3,500,000 = 3.5 \times 10^6$$

$$(b) 0.000789 = 7.89 \times 10^{-4}$$

$$(c) 300,000,000 = 3.0 \times 10^8 \text{ m/s}$$

Q6. How many sig figs in: (a) 0.0050 (b) 3.040 (c) 2500 (d) 1.00×10^3

$$(a) 0.0050 \rightarrow 2 \text{ (leading zeros not sig; trailing decimal zero IS)}$$

$$(b) 3.040 \rightarrow 4 \text{ (all digits including trailing zero after decimal)}$$

$$(c) 2500 \rightarrow \text{ambiguous (2,3 or 4)} \text{ — write in scientific notation to clarify}$$

$$(d) 1.00 \times 10^3 \rightarrow 3 \text{ (the three digits 1,0,0 are all significant)}$$

Q7. Calculate volume occupied by 3 moles of O_2 at STP.

$$\text{At STP: 1 mole of gas} = 22.4 \text{ L}$$

$$3 \text{ moles of } O_2 = 3 \times 22.4 = 67.2 \text{ L}$$

Q8. In $N_2 + 3H_2 \rightarrow 2NH_3$, find mass of NH_3 formed from 28 g of N_2 .

$$n(N_2) = 28/28 = 1 \text{ mol}$$

$$\text{From equation: 1 mol } N_2 \rightarrow 2 \text{ mol } NH_3$$

$$n(NH_3) = 2 \text{ mol}$$

$$\text{Mass} = 2 \times 17 = 34 \text{ g of } NH_3$$

Q9. Give two differences between homogeneous and heterogeneous mixtures.

Homogeneous: Uniform composition; single phase; components not visible separately. Example: salt water, air.

Heterogeneous: Non-uniform; two or more visible phases; components distinguishable. Example: sand in water, salad.

Q10. What does 1 u equal in grams?

1 u = $1/12 \times$ (mass of one ^{12}C atom)

$$1 \text{ u} = 1.66 \times 10^{-24} \text{ g}$$

This tiny value explains why we use relative atomic masses and moles instead of actual gram values.

Q11. Calculate formula mass of CaCO_3 . (Ca=40, C=12, O=16)

$$\text{CaCO}_3 = \text{Ca} + \text{C} + 3\text{O}$$

$$= 40 + 12 + 3(16)$$

$$= 40 + 12 + 48$$

$$= 100 \text{ g/mol} \text{ (important compound — chalk, limestone, antacids!)}$$

Q12. Differentiate between atom and molecule with examples.

Atom: Smallest particle of an element capable of chemical reaction. May or may not exist independently. Examples: Na, Fe, Cu (monoatomic).

Molecule: Two or more atoms bonded together. Examples: H_2O , O_2 , CO_2 . Note: H_2 , O_2 , N_2 are diatomic molecules of elements.

Q13. Given: 4g H_2 and 32g O_2 for $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. Find limiting reagent.

$$n(\text{H}_2) = 4/2 = 2 \text{ mol}$$

$$n(\text{O}_2) = 32/32 = 1 \text{ mol}$$

Divide by coefficients: H_2 : $2/2 = 1.0$; O_2 : $1/1 = 1.0$

Both equal \rightarrow **both consumed completely (exact stoichiometric ratio!)**

$$\text{H}_2\text{O formed} = 2 \text{ mol} \times 18 = 36 \text{ g}$$

Q14. Find mass of 3.011×10^{23} molecules of CO_2 .

$$n = 3.011 \times 10^{23} / 6.022 \times 10^{23} = 0.5 \text{ mol}$$

$$\text{Mass} = 0.5 \times 44 = 22 \text{ g of } \text{CO}_2$$

Q15. State two differences between elements and compounds.

Element: Made of one type of atom; cannot be decomposed by chemical methods; 118 known. Example: Fe, Au, O_2 .

Compound: Two or more elements in fixed ratio; can be decomposed chemically; properties differ from elements. Example: H_2O , NaCl.

Q16. What is the mass of 0.5 mol of H_2SO_4 ?

Molar mass of $\text{H}_2\text{SO}_4 = 2+32+64 = 98 \text{ g/mol}$

Mass = $n \times M = 0.5 \times 98 = 49 \text{ g}$

Q17. How many moles are in 1.204×10^{24} atoms of iron?

$n = \text{Number of atoms} / N_A$

$n = 1.204 \times 10^{24} / 6.022 \times 10^{23} = 2 \text{ moles of Fe}$

Q18. Actual yield = 8g, Theoretical yield = 10g. Calculate percentage yield.

% Yield = $(\text{Actual/Theoretical}) \times 100$

= $(8/10) \times 100$

= 80%

(20% product was lost due to side reactions, spillage, or incomplete reaction)

Q19. Calculate moles and number of molecules in 11g of CO_2 .

Molar mass $\text{CO}_2 = 44 \text{ g/mol}$

$n = 11/44 = 0.25 \text{ mol}$

Molecules = $0.25 \times 6.022 \times 10^{23} = 1.5055 \times 10^{23} \text{ molecules}$

Q20. What is the mole ratio in: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$? How is it used?

Mole ratio: $\text{H}_2 : \text{O}_2 : \text{H}_2\text{O} = 2 : 1 : 2$

Usage: If we have 4 mol H_2 , we need 2 mol O_2 and get 4 mol H_2O .

Mole ratios are the key to ALL stoichiometry calculations — always start with them!

SECTION 6: LONG ANSWER QUESTIONS

10 Questions | 5 marks each | Detailed answers for NEET/IIT-JEE/Board essays

Q1. Explain all five Laws of Chemical Combination in detail with examples and significance.

■ Background

Before Dalton's atomic theory, scientists discovered repeating patterns in chemical reactions. These patterns became the five Laws of Chemical Combination — the foundation of modern chemistry.

■ 1. Law of Conservation of Mass (Lavoisier, 1789)

"In any chemical reaction, total mass of reactants = total mass of products."

Example: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$. If $24\text{g Mg} + 32\text{g O}_2 \rightarrow 56\text{g MgO}$. Mass before = $24+32 = 56\text{g}$ = mass after ✓

■ 2. Law of Definite Proportions (Proust, 1799)

"A pure compound always has the same elements in the same fixed mass ratio."

Example: H_2O always has H:O = 1:8 by mass, whether from Ganges river or laboratory.

■ 3. Law of Multiple Proportions (Dalton, 1803)

"When two elements form more than one compound, masses of one element (for fixed mass of other) are in simple whole number ratios."

Example: $\text{CO} \rightarrow 16\text{g O per } 12\text{g C}$; $\text{CO}_2 \rightarrow 32\text{g O per } 12\text{g C}$. Ratio = 1:2 ✓

■ 4. Gay-Lussac's Law of Gaseous Volumes

"At constant T and P, volumes of gaseous reactants/products are in simple whole number ratios."

Example: $2 \text{ vol H}_2 + 1 \text{ vol O}_2 \rightarrow 2 \text{ vol H}_2\text{O}$. Ratio = 2:1:2.

■ 5. Law of Reciprocal Proportions (Richter)

"If A and B each combine with element C, then A and B combine with each other in the same ratio or a simple multiple."

■ Significance

These five laws collectively led John Dalton to formulate his Atomic Theory in 1808, which explained ALL of them using the concept of atoms.

Q2. Describe Dalton's Atomic Theory completely — postulates, explanations, and limitations.

■ Historical context

John Dalton (1766–1844), an English chemist and school teacher, proposed the first scientific Atomic Theory in 1808 based on experimental evidence from chemical laws.

■ Five Postulates

1. All matter consists of tiny, indivisible particles called ATOMS.
2. All atoms of a given element are identical in mass and chemical properties.
3. Atoms of different elements have different masses and properties.
4. Atoms combine in simple whole number ratios to form compounds (e.g., $\text{H}_2\text{O} = 2\text{H} + 1\text{O}$).
5. In chemical reactions, atoms are rearranged only — never created or destroyed.

■ How it explains chemical laws

- Explains **Conservation of Mass** (postulate 5 — atoms just rearrange, total count unchanged).
- Explains **Definite Proportions** (postulate 4 — fixed atomic ratio means fixed mass ratio).
- Explains **Multiple Proportions** (different but always whole number combinations).

■ Three Key Limitations

1. **Atoms ARE divisible:** J.J. Thomson (1897) discovered electrons inside atoms. Atoms contain protons, neutrons, and electrons.
2. **Isotopes exist:** Atoms of the same element can have different masses. C-12 and C-14 — same element, different neutron numbers.
3. **Non-stoichiometric compounds:** Some compounds like $\text{FeO}_{0.95}$ do not have simple whole-number ratios.

■ Legacy

Despite limitations, Postulate 5 — "atoms rearrange but are never created or destroyed" — remains the core principle of chemistry. Every balanced equation is based on this fundamental truth.

Q3. Explain the Mole Concept in detail. Cover: definition, Avogadro's number, molar mass, molar volume, and worked examples.

■ **Why we need the mole**

Atoms are incredibly small: H atom mass $\approx 1.66 \times 10^{-24}$ g. We cannot weigh or count individual atoms. We need a practical unit — just as we count eggs in "dozens" (12), we count atoms in "moles".

■ **Definition**

1 mole = 6.022×10^{23} particles (atoms, molecules, ions, electrons, etc.).

This number is called **Avogadro's Number (N_A)**.

■ **Avogadro's Number Scale**

6.022×10^{23} is enormous. If 1 mole of rice grains were spread over Earth's surface, they would form a layer several kilometres deep. Yet, just 18 g of water contains this many molecules!

■ **Molar Mass**

Molar mass = mass of 1 mole of substance (g/mol). Numerically equals atomic/molecular mass in "u".

Examples: $H_2O = 18$ g/mol; $Fe = 56$ g/mol; $CO_2 = 44$ g/mol.

Key formula: $n = w/M$

■ **Molar Volume**

At STP ($0^\circ C$, 1 atm): **1 mole of any ideal gas = 22.4 L**.

This is the same for every gas — 2g H_2 , 32g O_2 , 44g CO_2 all occupy 22.4 L at STP.

■ **Worked Examples**

Q: Molecules in 36g H_2O ?

$n = 36/18 = 2$ mol. Molecules = $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$.

Q: Mass of 3.011×10^{23} atoms of Fe?

$n = 3.011 \times 10^{23} / 6.022 \times 10^{23} = 0.5$ mol. Mass = $0.5 \times 56 = 28$ g.

Q4. Explain Stoichiometry with the ammonia synthesis reaction. Include mole ratios, mass calculations, and limiting reagent concept.

■ What is Stoichiometry?

From Greek: stoicheion (element) + metron (measure). It is the quantitative calculation of how much reactant is needed and how much product is formed in a chemical reaction.

■ Balanced Equation: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

Reading this equation:

- 1 mol N_2 + 3 mol H_2 \rightarrow 2 mol NH_3
- 28g N_2 + 6g H_2 \rightarrow 34g NH_3 (28+6=34: mass conserved ✓)
- Mole ratio: $\text{N}_2:\text{H}_2:\text{NH}_3 = 1:3:2$

■ 4-Step Problem Solving Method

Step 1: Balance the equation.

Step 2: Convert given mass \rightarrow moles ($n = w/M$).

Step 3: Use mole ratio from equation.

Step 4: Convert moles \rightarrow required unit (mass, volume, particles).

■ Worked Example

Q: Mass of NH_3 from 14g N_2 ?

$$n(\text{N}_2) = 14/28 = 0.5 \text{ mol}$$

$$\text{Ratio: } 1 \text{ mol } \text{N}_2 \rightarrow 2 \text{ mol } \text{NH}_3$$

$$n(\text{NH}_3) = 0.5 \times 2 = 1 \text{ mol}$$

$$\text{Mass} = 1 \times 17 = \mathbf{17 \text{ g } \text{NH}_3}$$

■ Applications

Stoichiometry is used in: industrial chemical production, pharmaceutical manufacturing, environmental monitoring, and food science.

Q5. Explain the concept of Limiting Reagent and Percentage Yield with detailed examples.

■ Real-life analogy

To make 10 cups of tea you need: milk (have enough for 20 cups) + tea leaves (enough for only 8 cups). You can make only 8 cups — because tea leaves run out first. Tea leaves = LIMITING REAGENT. Milk = EXCESS REAGENT.

■ Definition

The **limiting reagent** is the reactant completely consumed first in a reaction. It LIMITS how much product can form. The other reactant (in excess) is partially consumed — some remains unused.

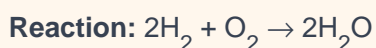
■ How to Identify Limiting Reagent

Step 1: Convert all masses to moles.

Step 2: Divide each by its stoichiometric coefficient.

Step 3: Smallest result = Limiting Reagent.

■ Detailed Example



Given: 6g H_2 and 40g O_2

$n(\text{H}_2) = 6/2 = 3$ mol. Divide by coeff: $3/2 = 1.5$

$n(\text{O}_2) = 40/32 = 1.25$ mol. Divide by coeff: $1.25/1 = 1.25$

Smaller value = $\text{O}_2 \rightarrow \text{O}_2$ is the Limiting Reagent!

H_2O formed = 1.25 mol \times 18 = 22.5g

Excess $\text{H}_2 = (3 - 2 \times 1.25) = 0.5$ mol = 1g unreacted

■ Percentage Yield

% Yield = (Actual Yield / Theoretical Yield) \times 100

If theoretical = 22.5g but actual = 20g: % yield = $(20/22.5) \times 100 = 88.9\%$

The 11.1% was lost to side reactions, spillage, or incomplete reaction.

Q6. Explain Significant Figures in detail — all rules, examples, and how to apply them in calculations.

■ Why significant figures?

Every measurement has uncertainty. A ruler measuring 5 cm may actually be 4.9–5.1 cm. Significant figures (sig figs) tell us the precision of our measurement and how many digits we can trust.

■ 5 Rules with Examples

Rule 1 — Non-zero digits: Always significant. 234 → 3; 9.87 → 3.

Rule 2 — Zeros between non-zeros: Significant. 205 → 3; 10.07 → 4.

Rule 3 — Leading zeros: NOT significant. 0.0045 → 2 (only 4 and 5 count).

Rule 4 — Trailing zeros after decimal: Significant. 2.500 → 4; 3.00 → 3.

Rule 5 — Trailing zeros in whole numbers: Ambiguous. 2500 could be 2/3/4. Use scientific notation: $2.5 \times 10^3 = 2$ sig figs; $2.500 \times 10^3 = 4$ sig figs.

■ In Calculations

Multiplication/Division: Answer has as many sig figs as the LEAST precise measurement.

Example: $2.5 \times 3.42 = 8.55 \rightarrow$ round to 2 sig figs = **8.6**

Addition/Subtraction: Round to the fewest DECIMAL PLACES.

Example: $12.11 + 18.0 + 1.013 = 31.123 \rightarrow$ round to 1 decimal = **31.1**

■ Importance for NEET/JEE

Sig fig questions appear regularly. Remember: precision cannot increase through calculation — your answer is only as precise as your least precise measurement.

Q7. Classify matter in detail: elements, compounds, homogeneous and heterogeneous mixtures. How are they separated?

■ Overview

All matter is either a Pure Substance or a Mixture.

■ Pure Substances

Elements: Made of one type of atom. 118 elements in Periodic Table. Cannot be broken down chemically.

Examples: Cu (copper), Fe (iron), O₂ (oxygen), C (carbon/diamond/graphite).

Compounds: Two or more elements in fixed ratio, chemically bonded. Completely new properties.

Examples: H₂O (water), NaCl (salt), HCl (acid), CaCO₃ (limestone).

Separation: requires chemical methods (electrolysis, thermal decomposition).

■ Mixtures

Homogeneous: Uniform composition; single visible phase; cannot distinguish components by eye.

Examples: Salt water, air (N₂+O₂+others), vinegar, brass alloy.

Separation: Distillation, evaporation, chromatography.

Heterogeneous: Non-uniform; different phases/components visible.

Examples: Sand in water, oil+water, granite (quartz+feldspar+mica), salad.

Separation: Filtration, decantation, magnetic separation, sedimentation.

■ Key Comparisons

Compound vs Mixture:

- Compound: fixed ratio, chemical combination, new properties, chemical separation.
- Mixture: variable ratio, physical combination, retains properties, physical separation.

Example question: Is river water a mixture or compound?

Answer: Mixture! Contains H₂O + dissolved salts + gases + microorganisms. Pure H₂O is a compound.

Q8. Explain atomic mass, molecular mass, formula mass, and molar mass with examples and calculations.

■ **Why relative atomic mass?**

Actual atomic masses are too small to work with directly ($H = 1.66 \times 10^{-24}$ g). Scientists defined relative atomic masses using carbon-12 as the standard.

■ **Atomic Mass**

Mass of an atom relative to $1/12$ the mass of ^{12}C . Unit: u (unified mass unit).

Standard: $^{12}\text{C} = \text{exactly } 12.000 \text{ u}$.

Key values to memorise: H=1, C=12, N=14, O=16, Na=23, S=32, Cl=35.5, Fe=56, Cu=63.5

■ **Molecular Mass**

Sum of atomic masses of all atoms in one molecule (for covalent compounds).

$$\text{H}_2\text{SO}_4 = 2(1) + 32 + 4(16) = 2 + 32 + 64 = \mathbf{98 \text{ u}}$$

$$\text{C}_6\text{H}_{12}\text{O}_6 = 6(12) + 12(1) + 6(16) = 72 + 12 + 96 = \mathbf{180 \text{ u}}$$

■ **Formula Mass**

For ionic compounds (no discrete molecules). Same calculation method, different name.

$$\text{NaCl} = 23 + 35.5 = 58.5 \text{ u}$$

$$\text{CaCO}_3 = 40 + 12 + 3(16) = 100 \text{ u}$$

■ **Molar Mass**

Numerical value same as atomic/molecular mass, but unit = g/mol.

$18 \text{ u} \rightarrow 18 \text{ g/mol}$ means: 18 grams of H_2O contains exactly 6.022×10^{23} molecules.

This connection between u and g/mol is a fundamental bridge between atomic and lab scales.

Q9. What is the historical development of atomic theory from Acharya Kanad to modern times?

■ **Ancient India — Acharya Kanad (~600 BCE)**

Acharya Kanad (Kashyap) proposed that all matter is made of tiny, indivisible particles called "**Parmanu**" (परमाणु) — meaning "that which cannot be cut further". This was philosophical reasoning without experiments, yet remarkably anticipates modern atomic theory by ~2400 years.

■ **Ancient Greece — Democritus (~460 BCE)**

Greek philosopher Democritus independently proposed "atomos" (indivisible). However, Aristotle rejected this, delaying Western atomic theory by ~2000 years.

■ **Antoine Lavoisier (1789)**

Established Law of Conservation of Mass through careful quantitative experiments — laying the experimental foundation.

■ **Joseph Proust (1799)**

Law of Definite Proportions — compounds always have fixed composition, suggesting discrete atomic combinations.

■ **John Dalton (1808)**

First scientific atomic theory — 5 postulates explaining all known chemical laws. First quantitative atomic model.

■ **J.J. Thomson (1897)**

Discovered the electron — proving atoms ARE divisible. Proposed "Plum Pudding" model.

■ **Rutherford (1911)**

Gold foil experiment → nucleus discovered. Planetary model of atom.

■ **Bohr (1913) → Modern Quantum Model**

Electron shells and orbitals. Schrodinger's quantum mechanical model (modern, most accurate).

■ **James Chadwick (1932)**

Discovered neutron — explained isotopes (Dalton's limitation resolved).

Q10. Write a comprehensive note on Stoichiometry. Include balanced equations, mole ratios, and a complete numerical problem.

■ Definition and Scope

Stoichiometry is the branch of chemistry dealing with quantitative relationships in chemical reactions. It allows chemists to predict: (1) how much reactant is needed, (2) how much product will form, (3) which reactant limits the reaction.

■ Interpreting a Balanced Equation

Consider: $2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$

This tells us: 2 mol Al reacts with 3 mol Cl_2 to give 2 mol AlCl_3 .

Mass: $54\text{g Al} + 213\text{g Cl}_2 \rightarrow 267\text{g AlCl}_3$ ($54+213=267$: conservation ✓).

■ 4-Step Method

Step 1: Write and balance the equation.

Step 2: Convert given quantities to moles ($n = w/M$).

Step 3: Use mole ratios from balanced equation.

Step 4: Convert to required units (mass, volume, particles).

■ Complete Numerical Problem

Problem: 5.4g of Al reacts with excess Cl_2 ($2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$). Find: (a) mass of AlCl_3 formed, (b) mass of Cl_2 consumed.

Solution:

Molar masses: $\text{Al}=27$, $\text{Cl}=35.5$, $\text{AlCl}_3=27+3(35.5)=133.5$, $\text{Cl}_2=71$

$n(\text{Al}) = 5.4/27 = 0.2 \text{ mol}$

(a) $2 \text{ mol Al} \rightarrow 2 \text{ mol AlCl}_3$; $n(\text{AlCl}_3) = 0.2 \text{ mol}$; Mass = $0.2 \times 133.5 = \mathbf{26.7g}$

(b) $2 \text{ mol Al} \rightarrow 3 \text{ mol Cl}_2$; $n(\text{Cl}_2) = 0.2 \times (3/2) = 0.3 \text{ mol}$; Mass = $0.3 \times 71 = \mathbf{21.3g}$

■ Industrial Importance

In the Haber process (making fertiliser), stoichiometry determines exact amounts of N_2 and H_2 to use. In pharmaceuticals, stoichiometry ensures drug purity and correct dosage. Even cooking recipes are a form of stoichiometry!

ALL THE BEST FOR YOUR EXAMS!

Chemistry Ch. 1 | Some Basic Concepts of Chemistry

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30 MCQs · 20 Fill Blanks · 10 Matching · 30 FAQs · 20 Short Q&A · 10 Long Answers

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