ATOMS AND MOLECULES

CBSE CLASS 9 SCIENCE (CHEMISTRY) • CHAPTER 3 • COMPLETE HIGH-YIELD QUESTION BANK & NOTES

Chapter Overview & Strategic Matrix

Aspect	Details
Weightage	10–12% in Annual CBSE Examinations (High Weightage Core Chapter)
Key Concepts	Laws of Chemical Combination, Dalton's Atomic Theory, Relative Atomic Mass Units, Criss-Cross Chemical Formula Writing, Mole Concept, Stoichiometric Mass-Particle Conversions
High Probability Map	Mole concept mathematical conversions, writing polyatomic ionic formulae, laws verification numericals (3–5 Marks)
Medium Probability Map	Dalton's theory drawbacks, relative mass scales, classification of ions (2–3 Marks)

Core Chapter Lexicon

Term	Strict Chemical Definition
Atom	The smallest defining unit of an element that retains its elemental identity and takes part directly in a chemical reaction.
Molecule	The smallest particle of an element or compound capable of completely independent stable existence under normal environmental conditions.
Atomic Mass	The mass of a single atom of an element evaluated relative to exactly 1/12th the mass of a isolated Carbon-12 atom, expressed in unified units (u).
Molecular Mass	The sum total of the relative atomic masses of all constituent atoms present within a single molecule of a substance.
Formula Unit Mass	The sum total of atomic masses of all atoms present within a formula unit of an ionic crystal lattice (used where discrete molecules do not exist).
Mole	The exact chemical amount of a substance that contains precisely 6.022×10^{23} fundamental elementary particles.
Molar Mass	The absolute mass of exactly one mole of a substance, expressed in grams per mole (g/mol). It is numerically equivalent to the atomic/molecular mass.
Valency	The combining capacity of an atom or group of atoms, dictated by its electronic structure.

Fundamental Kinematic & Stoichiometric Mathematical Formulations

- **Number of Moles (n) from Given Mass (m):**

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$$n =  rac{ ext{Given Mass } (m)}{ ext{Molar Mass } (M)}$$
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- **Number of Moles (n) from Particle Count (N):**

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$$n =  rac{ ext{Number of Particles } (N)}{ ext{Avogadro's Number } (N_a)}$$
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- **Total Particle Count (N):**

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$$N = n imes N_a =  rac{m}{M} imes 6.022 imes 10^{23}$$
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- **Mass Percentage of an Element:**

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$$ ext{Mass } \% =  rac{ ext{Total Mass of Element in Compound}}{ ext{Molar Mass of Compound}} imes 100$$
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Constants to Remember: Avogadro's Number (N_a) = $6.022 imes 10^{23} ext{ particles/mol}$. Unified Mass Unit ($1 ext{ u}$) = $1.6605 imes 10^{-24} ext{ g}$.

Standard Relative Atomic Masses (NCERT Benchmark Values)

Element	Symbol	Atomic Mass (u)
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24

Element	Symbol	Atomic Mass (u)
Aluminium	Al	27
Phosphorus	P	31
Sulphur	S	32
Chlorine	Cl	35.5
Calcium	Ca	40
Iron	Fe	56

SECTION A: Top 50 Most Important Questions with Answers

A. Very Short Answer Questions (1 Mark)

Q1. State the law of conservation of mass.

> Mass can neither be created nor destroyed during the course of a chemical reaction.

Q2. Who proposed the law of constant proportions?

> Joseph Louis Proust in 1779.

Q3. What is the atomic mass of carbon?

> 12 u (relative to the Carbon-12 standard isotope).

Q4. Define atomicity. Give an example of a diatomic molecule.

> Atomicity is the total number of atoms present within a single molecule of an element. Examples of diatomic molecules include H_2, O_2, and N_2.

Q5. What is the valency of oxygen?

> 2 (since it requires 2 electrons to complete its outer valence octet).

Q6. What is the chemical formula of sodium chloride?

> NaCl.

Q7. Name the scientist who gave the atomic theory of matter.

> John Dalton in 1808.

Q8. What is Avogadro's number?

> The constant value of 6.022×10^{23}, which represents the absolute number of elementary entities present in one mole of any substance.

Q9. Calculate the number of moles in 12 g of carbon.

> $$\text{Number of moles } (n) = \frac{\text{Given Mass } (m)}{\text{Molar Mass } (M)} = \frac{12 \text{ g}}{12 \text{ g/mol}} = 1 \text{ mole}$$

Q10. What is the mass of one mole of water?

> 18 grams ($\text{H}_2\text{O} = 2 \times 1 + 16 = 18 \text{ g/mol}$).

Q11. Give the formula of calcium hydroxide.

> Ca(OH)_2.

Q12. What is an ion? Give one example of a cation.

> An ion is an atom or molecule that carries a net electrical charge due to the loss or gain of electrons. Example of a cation: Na^+.

Q13. What is the molecular mass of $ ext{CO}_2$?

> 44 u ($ ext{C}=12 + 2 imes ext{O}(16) = 12 + 32 = 44 ext{ u}$).

Q14. How many atoms are present in one molecule of $ ext{H}_2 ext{SO}_4$?

> 7 atoms ($2 ext{ atoms of H} + 1 ext{ atom of S} + 4 ext{ atoms of O} = 7$).

Q15. What is the formula unit mass of $ ext{NaCl}$?

> 58.5 u ($ ext{Na}=23 + ext{Cl}=35.5 = 58.5 ext{ u}$).

B. Short Answer Questions (2–3 Marks)

Q1. State the law of constant proportions. Explain with an example.

> The law of constant proportions states that in a pure chemical substance, the constituent elements are always combined in a fixed and definite proportion by mass, regardless of the source or method of preparation.
> *Example:* Pure water ($ ext{H}_2 ext{O}$) always contains Hydrogen and Oxygen combined in a strict mass ratio of $1:8$. If $9 ext{ g}$ of water is decomposed by electrolysis, it will always yield exactly $1 ext{ g}$ of Hydrogen gas and $8 ext{ g}$ of Oxygen gas.

Q2. Give two central postulates of Dalton's atomic theory.

> 1. All matter is composed of microscopic, indivisible particles called atoms.
> 2. Atoms can neither be created nor destroyed during any chemical reaction.
> 3. Atoms of a given element are identical in mass and chemical properties, while atoms of different elements differ.

Q3. Distinguish clearly between an atom and a molecule.

Criteria	Atom	Molecule
Definition	The smallest component particle of an element.	The smallest independent particle of an element or compound.
Stability	Highly reactive; rarely exists independently in nature (except noble gases).	Stable structure; capable of completely independent natural existence.
Examples	$ ext{H}$, $ ext{O}$, $ ext{Na}$, $ ext{Fe}$	$ ext{H}_2$, $ ext{O}_2$, $ ext{H}_2 ext{O}$, $ ext{CO}_2$

Q4. What is atomic mass? Why is it called relative atomic mass?

> Atomic mass is the mass of an individual atom expressed in unified atomic mass units (u). Because weighing a single atom is practically impossible, its mass is measured on a relative scale compared to a standard reference. Since 1961, the standard reference has been the Carbon-12 isotope, which is assigned a mass of exactly 12 u. Therefore, it is called relative atomic mass because it shows how many times an atom of an element is heavier than $ rac{1}{12} ext{th}$ the mass of a single Carbon-12 atom.

Q5. Calculate the molecular mass of $ ext{H}_2 ext{O}$, $ ext{CO}_2$, and $ ext{CH}_4$.

> - **$ ext{H}_2 ext{O}$:** $(2 imes 1 ext{ u}) + (1 imes 16 ext{ u}) = 2 + 16 = 18 ext{ u}$
> - **$ ext{CO}_2$:** $(1 imes 12 ext{ u}) + (2 imes 16 ext{ u}) = 12 + 32 = 44 ext{ u}$
> - **$ ext{CH}_4$:** $(1 imes 12 ext{ u}) + (4 imes 1 ext{ u}) = 12 + 4 = 16 ext{ u}$

Q6. Define valency. Give the valency of (a) Sodium, (b) Magnesium, (c) Chlorine.

> Valency is the combining capacity of an atom, determined by the number of valence electrons it must gain, lose, or share to achieve a stable octet configuration.
> (a) **Sodium ($ ext{Na}$):** Configuration is $2, 8, 1$. It loses 1 electron, so its valency is **1**.
> (b) **Magnesium ($ ext{Mg}$):** Configuration is $2, 8, 2$. It loses 2 electrons, so its valency is **2**.
> (c) **Chlorine ($ ext{Cl}$):** Configuration is $2, 8, 7$. It gains 1 electron ($8 - 7$), so its valency is **1**.

Q7. Derive the chemical formulae of: (a) Aluminium oxide, (b) Sodium carbonate, (c) Calcium phosphate.

> Using the valence criss-cross method:
> (a) **Aluminium oxide:** Ions $ ext{Al}^{3+}$ and $ ext{O}^{2-} ightarrow$ Criss-cross subscripts $ ightarrow$ **$ ext{Al}_2 ext{O}_3$**
> (b) **Sodium carbonate:** Ions $ ext{Na}^{+}$ and $ ext{CO}_3^{2-} ightarrow$ Criss-cross subscripts $ ightarrow$ **$ ext{Na}_2 ext{CO}_3$**
> (c) **Calcium phosphate:** Ions $ ext{Ca}^{2+}$ and $ ext{PO}_4^{3-} ightarrow$ Criss-cross subscripts $ ightarrow$ **$ ext{Ca}_3(ext{PO}_4)_2$**

Q8. Differentiate between a cation and an anion with clear examples.

Cation	Anion
Carries a net positive electrical charge.	Carries a net negative electrical charge.
Formed when a neutral atom loses valence electrons.	Formed when a neutral atom gains extra electrons.
Typically formed by metal atoms (e.g., $ ext{Na}^+$, $ ext{Ca}^{2+}$, $ ext{NH}_4^+$).	Typically formed by non-metal atoms (e.g., $ ext{Cl}^-$, $ ext{O}^{2-}$, $ ext{SO}_4^{2-}$).

Q9. What is the mole concept? Why is it important?

> The mole concept provides a mathematical link between the microscopic scale of atoms or molecules and the macroscopic scale of grams used in laboratories. It defines a mole as the specific amount of a substance that contains exactly $6.022 imes 10^{23}$ particles. This concept is essential because it allows scientists to "count" atoms or molecules by weighing out a specific mass of a substance.

Q10. Calculate the number of moles in 22 g of carbon dioxide.

> $$ ext{Molar mass of } ext{CO}_2 = 12 + (2 imes 16) = 44 ext{ g/mol}$$ $$ ext{Number of moles } (n) = rac{ ext{Given Mass } (m)}{ ext{Molar Mass } (M)} = rac{22 ext{ g}}{44 ext{ g/mol}} = 0.5 ext{ mole}$$

Q11. Calculate the mass of 0.5 mole of nitrogen gas ($ ext{N}_2$).

> $$ ext{Molar mass of diatomic } ext{N}_2 = 2 imes 14 = 28 ext{ g/mol}$$ $$ ext{Mass } (m) = ext{Moles } (n) imes ext{Molar Mass } (M) = 0.5 imes 28 ext{ g} = 14 ext{ grams}$$

Q12. How many molecules are present in 36 g of water?

> $$ ext{Molar mass of } ext{H}_2 ext{O} = 18 ext{ g/mol} \implies n = rac{36 ext{ g}}{18 ext{ g/mol}} = 2 ext{ moles}$$ $$ ext{Molecules } (N) = n imes N_a = 2 imes 6.022 imes 10^{23} = 1.2044 imes 10^{24} ext{ molecules}$$

Q13. What is the formula unit mass of $ ext{CaCl}_2$ and $ ext{Na}_2 ext{CO}_3$?

> • **$ ext{CaCl}_2$:** $ ext{Ca}(40) + 2 imes ext{Cl}(35.5) = 40 + 71 = 111 ext{ u}$
> • **$ ext{Na}_2 ext{CO}_3$:** $2 imes ext{Na}(23) + ext{C}(12) + 3 imes ext{O}(16) = 46 + 12 + 48 = 106 ext{ u}$

Q14. Write the molecular formula and state the atomicity of: (a) Oxygen, (b) Phosphorus, (c) Sulphur.

> (a) **Oxygen:** Formula is $ ext{O}_2$; Atomicity = 2 (Diatomic)
> (b) **Phosphorus:** Formula is $ ext{P}_4$; Atomicity = 4 (Tetratomic)
> (c) **Sulphur:** Formula is $ ext{S}_8$; Atomicity = 8 (Polyatomic)

Q15. State two distinct limitations of Dalton's atomic theory.

> 1. Dalton stated that atoms are indivisible. This was disproved by the discovery of subatomic particles like protons, electrons, and neutrons.
> 2. He claimed that all atoms of a given element have identical masses. This was disproved by the discovery of **isotopes** (atoms of the same element with different masses, like Chlorine-35 and Chlorine-37).

C. Long Answer Questions (5 Marks)

Q1. State and explain the central laws of chemical combination with illustrative examples.

Law Name & Proposer	Core Scientific Statement	Illustrative Example
Law of Conservation of Mass (Antoine Lavoisier)	The total mass of substances remains completely constant during a chemical change. Mass can neither be created nor destroyed in a reaction.	Heating $10 ext{ g}$ of $ ext{CaCO}_3$ yields exactly $5.6 ext{ g}$ of solid $ ext{CaO}$ and $4.4 ext{ g}$ of $ ext{CO}_2$ gas ($5.6 + 4.4 = 10 ext{ g}$).
Law of Constant Proportions (Joseph Proust)	A pure chemical compound always contains its constituent elements combined in an absolute, fixed proportion by mass, regardless of its source.	Pure water ($ ext{H}_2 ext{O}$) always features Hydrogen and Oxygen combined in a strict $1:8$ mass ratio, whether sourced from rain, a river, or a lab.
Law of Multiple Proportions (John Dalton)	When two elements combine to form more than one compound, the different masses of one element that combine with a fixed mass of the other element form a simple, whole-number ratio.	Carbon and Oxygen form $ ext{CO}$ ($12:16$) and $ ext{CO}_2$ ($12:32$). For a fixed mass of Carbon ($12 ext{ g}$), the mass ratio of Oxygen is $16:32$, which simplifies to a whole-number ratio of $1:2$.

Q2. Differentiate clearly between atomic mass, molecular mass, and formula unit mass with representative calculations.

1. Atomic Mass: The relative mass of an individual atom compared to the Carbon-12 standard unit reference ($1/12 ext{th}$ of a C-12 atom). It represents a single isolated atom. *Example:* The relative atomic mass of Sodium ($ ext{Na}$) is $23 ext{ u}$.

2. Molecular Mass: The sum total of the individual relative atomic masses of all atoms present in a discrete molecule. This term is used for covalent compounds that exist as separate molecules.
Example: Calculation for Nitric Acid ($ ext{HNO}_3$): $$ ext{Molecular Mass} = ext{Mass of H}(1) + ext{Mass of N}(14) + 3 imes ext{Mass of O}(16) = 1 + 14 + 48 = 63 ext{ u}$$

3. Formula Unit Mass: Calculated in the exact same way as molecular mass, but used specifically for compounds that form three-dimensional ionic crystal lattices rather than discrete molecules.
Example: Calculation for Sodium Carbonate ($ ext{Na}_2 ext{CO}_3$): $$ ext{Formula Unit Mass} = 2 imes ext{Na}(23) + ext{C}(12) + 3 imes ext{O}(16) = 46 + 12 + 48 = 106 ext{ u}$$

Q3. Write the chemical formulae of the following step-by-step using the criss-cross method: (a) Magnesium chloride, (b) Aluminium sulphate, (c) Calcium carbonate, (d) Sodium phosphate, (e) Ammonium sulphate.

Compound Name	Constituent Ions & Charges	Valence Criss-Cross Step	Final Formula
Magnesium chloride	$ ext{Mg}^{2+}$ and $ ext{Cl}^{-}$	$ ext{Mg}_1 \leftarrow ext{Cl}_2$	**$ ext{MgCl}_2$**
Aluminium sulphate	$ ext{Al}^{3+}$ and $ ext{SO}_4^{2-}$	$ ext{Al}_2 \leftarrow (ext{SO}_4)_3$	**$ ext{Al}_2(ext{SO}_4)_3$**
Calcium carbonate	$ ext{Ca}^{2+}$ and $ ext{CO}_3^{2-}$	$ ext{Ca}_2(ext{CO}_3)_2 ightarrow$ simplify ratio	**$ ext{CaCO}_3$**
Sodium phosphate	$ ext{Na}^{+}$ and $ ext{PO}_4^{3-}$	$ ext{Na}_3 \leftarrow (ext{PO}_4)_1$	**$ ext{Na}_3 ext{PO}_4$**
Ammonium sulphate	$ ext{NH}_4^{+}$ and $ ext{SO}_4^{2-}$	$(ext{NH}_4)_2 \leftarrow (ext{SO}_4)_1$	**$(ext{NH}_4)_2 ext{SO}_4$**

Q4. Explain how valency is determined from an element's electronic configuration. Provide examples for both metals and non-metals.

Valency is determined by the number of valence electrons present in the outermost shell of an atom. It is calculated using two standard rules based on the octet requirements:

- **For Metals (Valence Electrons ≤ 4):** The valency is exactly equal to the total number of outermost valence electrons, which the atom tends to lose.
  - *Example (Sodium - $ ext{Na}$):* Configuration is $2, 8, 1$. Outermost electron = 1. Valency = **1**.
  - *Example (Magnesium - $ ext{Mg}$):* Configuration is $2, 8, 2$. Outermost electrons = 2. Valency = **2**.
- **For Non-Metals (Valence Electrons > 4):** The valency is calculated by subtracting the number of valence electrons from 8, representing the electrons the atom needs to gain or share.
  - *Example (Oxygen - $ ext{O}$):* Configuration is $2, 6$. Valence electrons = 6. Valency = $8 - 6 =$ **2**.
  - *Example (Chlorine - $ ext{Cl}$):* Configuration is $2, 8, 7$. Valence electrons = 7. Valency = $8 - 7 =$ **1**.

SECTION B: Important Numerical Questions (25 Most Expected)

Numerical 1: Calculate the number of moles in 52 g of helium (He).

> Molar mass of $ ext{He} = 4 ext{ g/mol}$.
> $$ ext{Moles } (n) = rac{m}{M} = rac{52 ext{ g}}{4 ext{ g/mol}} = 13 ext{ moles}$$

Numerical 2: Calculate the mass of 2.5 moles of carbon dioxide ($ ext{CO}_2$).

> Molar mass of $ ext{CO}_2 = 44 ext{ g/mol}$.
> $$ ext{Mass } (m) = n imes M = 2.5 imes 44 ext{ g} = 110 ext{ grams}$$

Numerical 3: Calculate the total number of water molecules present in 90 g of water ($ ext{H}_2 ext{O}$).

> Molar mass of $ ext{H}_2 ext{O} = 18 ext{ g/mol}$.
> $$ ext{Moles } (n) = rac{90}{18} = 5 ext{ moles}$$ $$ ext{Molecules } (N) = n imes N_a = 5 imes 6.022 imes 10^{23} = 3.011 imes 10^{24} ext{ molecules}$$

Numerical 4: Calculate the number of atoms present in 0.2 mole of carbon.

> $$ ext{Atoms } (N) = n imes N_a = 0.2 imes 6.022 imes 10^{23} = 1.2044 imes 10^{23} ext{ atoms}$$

Numerical 5: Calculate the absolute mass of a single individual atom of carbon in grams.

> One mole of Carbon contains $6.022 imes 10^{23}$ atoms and weighs $12 ext{ grams}$.
> $$ ext{Mass of 1 atom} = rac{ ext{Molar Mass}}{ ext{Avogadro's Number}} = rac{12 ext{ g}}{6.022 imes 10^{23}} = 1.992 imes 10^{-23} ext{ grams}$$

Numerical 6: Calculate the number of moles in 22 g of $ ext{CO}_2$.

> Molar mass of $ ext{CO}_2 = 44 ext{ g/mol}$.
> $$n = rac{m}{M} = rac{22 ext{ g}}{44 ext{ g/mol}} = 0.5 ext{ mole}$$

Numerical 7: Calculate the mass of $1.2044 imes 10^{24}$ molecules of water.

> $$n = rac{N}{N_a} = rac{1.2044 imes 10^{24}}{6.022 imes 10^{23}} = 2 ext{ moles}$$ $$ ext{Mass } (m) = n imes M = 2 imes 18 ext{ g} = 36 ext{ grams}$$

Numerical 8: Calculate the total number of oxygen atoms present in 88 g of $ ext{CO}_2$.

> $$ ext{Molar mass of } ext{CO}_2 = 44 ext{ g/mol} \implies n = rac{88}{44} = 2 ext{ moles of } ext{CO}_2$$ $$ ext{Since 1 molecule of } ext{CO}_2 ext{ contains 2 Oxygen atoms, total moles of O atoms} = 2 imes 2 = 4 ext{ moles}$$ $$ ext{Total Oxygen Atoms} = 4 imes 6.022 imes 10^{23} = 2.4088 imes 10^{24} ext{ atoms}$$

Numerical 9: Calculate the formula unit mass of calcium carbonate ($ ext{CaCO}_3$).

> $$ ext{Mass} = ext{Ca}(40) + ext{C}(12) + 3 imes ext{O}(16) = 40 + 12 + 48 = 100 ext{ u}$$

Numerical 10: Calculate the mass percentage of nitrogen in ammonia ($ ext{NH}_3$).

> $$ ext{Molar mass of } ext{NH}_3 = 14 + (3 imes 1) = 17 ext{ g/mol}$$ $$ ext{Mass } \% ext{ of N} = rac{ ext{Mass of Nitrogen}}{ ext{Total Molar Mass}} imes 100 = \left(rac{14}{17} ight) imes 100 = 82.35\%$$

Numerical 11: Find the mass of 10 moles of sodium sulphite ($ ext{Na}_2 ext{SO}_3$).

> $$ ext{Molar mass of } ext{Na}_2 ext{SO}_3 = (2 imes 23) + 32 + (3 imes 16) = 46 + 32 + 48 = 126 ext{ g/mol}$$ $$ ext{Mass } (m) = n imes M = 10 imes 126 = 1260 ext{ grams}$$

Numerical 12: How many moles are present in $3.011 imes 10^{23}$ molecules of $ ext{O}_2$?

> $$n = rac{N}{N_a} = rac{3.011 imes 10^{23}}{6.022 imes 10^{23}} = 0.5 ext{ mole}$$

Numerical 13: Calculate the mass of 0.5 mole of nitrogen atoms ($ ext{N}$).

> $$ ext{Mass} = n imes ext{Atomic Mass} = 0.5 imes 14 = 7 ext{ grams}$$

Numerical 14: Calculate the mass of 0.5 mole of nitrogen molecules ($ ext{N}_2$).

> $$ ext{Mass} = n imes ext{Molar Mass of } ext{N}_2 = 0.5 imes 28 = 14 ext{ grams}$$

Numerical 15: How many molecules are present in 34 g of ammonia ($ ext{NH}_3$)?

> $$ ext{Molar mass of } ext{NH}_3 = 17 ext{ g/mol} \implies n = rac{34}{17} = 2 ext{ moles}$$ $$ ext{Molecules} = 2 imes 6.022 imes 10^{23} = 1.2044 imes 10^{24} ext{ molecules}$$

Numerical 16: Calculate the total number of individual hydrogen atoms in 36 g of water.

> $$ ext{Moles of water } (n) = rac{36}{18} = 2 ext{ moles of } ext{H}_2 ext{O}$$ $$ ext{Each molecule contains 2 H atoms} \implies ext{Total moles of H atoms} = 2 imes 2 = 4 ext{ moles}$$ $$ ext{Total Hydrogen Atoms} = 4 imes 6.022 imes 10^{23} = 2.4088 imes 10^{24} ext{ atoms}$$

Numerical 17: Find the absolute molecular mass of Sulphuric Acid ($ ext{H}_2 ext{SO}_4$).

> $$ ext{Mass} = (2 imes 1) + 32 + (4 imes 16) = 2 + 32 + 64 = 98 ext{ u}$$

Numerical 18: Which sample contains a larger count of atoms: 100 g of metallic sodium ($ ext{Na}$) or 100 g of pure solid iron ($ ext{Fe}$)?

> $$ ext{Moles of Na} = rac{100 ext{ g}}{23 ext{ g/mol}} = 4.35 ext{ moles} \quad ext{Moles of Fe} = rac{100 ext{ g}}{56 ext{ g/mol}} = 1.79 ext{ moles}$$ Since Sodium ($ ext{Na}$) has a lower atomic mass, $100 ext{ g}$ of sodium contains significantly more moles and therefore a larger total count of atoms than iron.

Numerical 19: Calculate the absolute mass in grams of a single isolated molecule of water.

> $$ ext{Mass of } 1 ext{ molecule} = rac{ ext{Molar Mass of } ext{H}_2 ext{O}}{ ext{Avogadro's Number}} = rac{18 ext{ g}}{6.022 imes 10^{23}} = 2.989 imes 10^{-23} ext{ grams}$$

Numerical 20: An unknown sample features a chemical weight composition profile of 40% Carbon, 6.67% Hydrogen, and 53.33% Oxygen. Find its empirical formula.

> Calculate the atomic mole ratios:
> • $ ext{Carbon (C)} = rac{40}{12} = 3.33 \implies ext{Ratio} = rac{3.33}{3.33} = 1$
> • $ ext{Hydrogen (H)} = rac{6.67}{1} = 6.67 \implies ext{Ratio} = rac{6.67}{3.33} = 2$
> • $ ext{Oxygen (O)} = rac{53.33}{16} = 3.33 \implies ext{Ratio} = rac{3.33}{3.33} = 1$
> The empirical formula is **$ ext{CH}_2 ext{O}$**.

Numerical 21: Calculate the number of moles in $3.011 imes 10^{22}$ atoms of carbon.

> $$n = rac{N}{N_a} = rac{3.011 imes 10^{22}}{6.022 imes 10^{23}} = 0.05 ext{ mole}$$

Numerical 22: Find the mass of 0.25 mole of oxygen gas ($ ext{O}_2$).

> $$m = n imes M = 0.25 imes 32 ext{ g} = 8 ext{ grams}$$

Numerical 23: Calculate the number of molecules present in 0.25 mole of oxygen gas.

> $$N = n imes N_a = 0.25 imes 6.022 imes 10^{23} = 1.5055 imes 10^{23} ext{ molecules}$$

Numerical 24: Calculate the mass of $3.011 imes 10^{23}$ atoms of calcium ($ ext{Ca}$).

> $$n = rac{3.011 imes 10^{23}}{6.022 imes 10^{23}} = 0.5 ext{ mole} \implies m = 0.5 imes 40 ext{ g} = 20 ext{ grams}$$

Numerical 25: Calculate the number of molecules in 8 g of oxygen gas ($ ext{O}_2$).

> $$n = rac{m}{M} = rac{8 ext{ g}}{32 ext{ g/mol}} = 0.25 ext{ mole}$$ $$ ext{Molecules} = 0.25 imes 6.022 imes 10^{23} = 1.5055 imes 10^{23} ext{ molecules}$$

SECTION C: NCERT Activity-Based Questions

Activity 3.1: Mass Conservation Verification Protocol

Prepare separate solutions of Barium Chloride (BaCl_2) and Sodium Sulphate (Na_2SO_4). Suspend the BaCl_2 solution in an ignition tube inside a sealed Erlenmeyer flask containing the Na_2SO_4 solution. Weigh the entire apparatus. Invert the flask to mix the solutions.
Core Observation & Deduction: A white precipitate of Barium Sulphate (BaSO_4) forms instantly, showing a chemical reaction has occurred. Re-weighing the flask shows that the total mass remains completely unchanged. This experiment directly verifies the **Law of Conservation of Mass**.

Activity 3.2: Electrolytic Proportions Checklist

Perform the electrolysis of distilled water samples collected from different independent municipal sources.
Core Observation & Deduction: Regardless of the water source, the volume of hydrogen gas collected is always exactly twice the volume of oxygen gas. When evaluated by weight, the elements are always present in a fixed mass ratio of $1:8$. This directly verifies the **Law of Constant Proportions**.

SECTION D: Case Study Questions (CBSE-Style)

Case Study: Stoichiometric Evaluation of Closed Reactor Masses

An experimental chemistry student burns a $2.43 ext{ g}$ sample of clean Magnesium ribbon ($ ext{Mg}$) inside a sealed, oxygen-filled combustion container. After the reaction is complete, a pure white powder called Magnesium Oxide ($ ext{MgO}$) is formed, weighing exactly $4.03 ext{ g}$. The balanced reaction equation is written as: $2 ext{Mg} + ext{O}_2 ightarrow 2 ext{MgO}$.

Q1. Calculate the exact mass of gaseous oxygen consumed during this combustion reaction.

> **Ans:** According to the Law of Conservation of Mass, the mass of the reactants must equal the mass of the products: $$ ext{Mass of Oxygen } (m) = ext{Mass of MgO Product} - ext{Mass of Mg Reactant}$$ $$ ext{Mass of Oxygen } (m) = 4.03 ext{ g} - 2.43 ext{ g} = 1.60 ext{ grams}$$

Q2. Calculate the fundamental whole-number mass combining ratio of Magnesium to Oxygen in this compound.

> **Ans:** The mass combining ratio is calculated as: $$ ext{Ratio } ext{Mg} : ext{O} = rac{2.43 ext{ g}}{1.60 ext{ g}} = rac{24.3}{16} pprox 3:2$$ This fixed whole-number ratio verifies the Law of Constant Proportions.

Q3. How many moles of Magnesium Oxide ($ ext{MgO}$) were produced? (Atomic masses: $ ext{Mg}=24$, $ ext{O}=16$).

> **Ans:** $$ ext{Molar mass of } ext{MgO} = 24 + 16 = 40 ext{ g/mol}$$ $$ ext{Number of moles } (n) = rac{ ext{Given Mass}}{ ext{Molar Mass}} = rac{4.03 ext{ g}}{40 ext{ g/mol}} pprox 0.1 ext{ mole}$$

SECTION E: Common Mistakes Students Make & How to Correct Them

1. Mistake: Writing relative atomic mass values with the wrong units or as plain numbers.
Correction: Use unified atomic mass units (**u**) when representing a single atom or molecule, and grams per mole (**g/mol**) when writing molar mass values for stoichiometry.

2. Mistake: Forgetting to multiply by the atomicity subscript when converting molecular gas masses.
Correction: Remember that 1 mole of Nitrogen gas ($ ext{N}_2$) weighs $28 ext{ g}$ ($2 imes 14$), not $14 ext{ g}$. Always check the formula subscript before calculating molar masses.

3. Mistake: Swapping the charge meanings of cations and anions during formula writing.
Correction: Remember that **Cations** are positively charged ions ($+$) formed by losing electrons, while **Anions** are negatively charged ions ($-$) formed by gaining electrons.

SECTION F: One-Page Quick Revision Sheet

Essential Polyatomic Ions

Ion Name	Formula & Charge
Ammonium	NH_4^+
Hydroxide	OH^-
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Sulphate	SO_4^{2-}
Phosphate	PO_4^{3-}

Atomicity Framework Reference

- **Monoatomic (1):** He, Ne, Ar, Kr (Noble Gases)
- **Diatomic (2):** H_2, O_2, N_2, Cl_2
- **Triatomic (3):** O_3 (Ozone)
- **Tetratomic (4):** P_4 (Phosphorus)
- **Polyatomic (8):** S_8 (Sulphur ring structures)

High-Yield Exam Reminder Checklist: (1) Always balance the net ionic charges when using the criss-cross method; the positive charges must exactly cancel out the negative charges. (2) Avogadro's constant is always written with a positive exponent: 6.022×10^{23}. (3) Relative atomic mass is always measured against the Carbon-12 standard, where 1 u is exactly $\frac{1}{12}\text{th}$ the mass of a single C-12 atom.