Chapter 3 – Atoms and Molecules (Premium & Detailed Notes)

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1. Background – Why Study Atoms & Molecules?

- **Atoms and molecules are the basic building blocks of all matter – everything around us, from water and air to metals and living beings, is made of them.**

- **Studying them helps us understand how substances are formed and why they show different properties (like why salt dissolves in water but oil does not).**

- **It explains everyday changes such as melting of ice, rusting of iron, or burning of fuel.**

- **This knowledge forms the foundation of science and technology – from creating new medicines and materials to innovations in fields like nanotechnology.**

- **Most importantly, it connects the invisible microscopic world with the visible world we experience, making science meaningful and practical.**

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2. Early Theories of Matter

1. ***Maharishi Kanad (6th century BCE)***
    - ***Proposed: Parmanu = smallest indivisible particle of matter.***
        - ***Believed that combinations of these parmanus form different substances.***
2. ***Democritus (460–370 BCE)***
    - ***Coined the term atomos meaning “indivisible.”***

3. *Suggested that atoms differ in shape, size, and arrangement.*

3. *Dalton's Atomic Theory (1803)*

John Dalton gave a scientific basis to the idea of atoms.

Postulates:

1. ***Matter consists of indivisible atoms.***
2. ***All atoms of an element are identical in mass and properties.***
3. ***Atoms of different elements have different masses and properties.***
4. ***Atoms combine in simple whole-number ratios to form compounds.***
5. ***Atoms can neither be created nor destroyed in a chemical reaction.***

Limitations:

- ***Atoms are divisible (discovery of subatomic particles).***
- ***Isotopes (same element, different mass) and isobars (different element, same mass) exist — Dalton's theory didn't explain them.***

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4. Laws of Chemical Combination

Chemists noticed that when elements react with each other to form compounds, they always follow certain rules. These rules are called the Laws of Chemical Combination. In Class 9, we mainly study two important laws:

1. Law of Conservation of Mass

- ***Proposed by: Antoine Lavoisier in 1774, also known as the "Father of Modern Chemistry."***

- ***Statement: "Mass can neither be created nor destroyed in a chemical reaction."*** ***Meaning:***
- ***In any chemical change, atoms are simply rearranged.***
- ***No atom is lost and no new atom is created, so the total mass before reaction = total mass after reaction.***
- ***This law shows that chemical reactions are like "reshuffling" of atoms, not creation or destruction.***

Example:

- ***Reaction of barium chloride and sodium sulfate:***

  BaCl2+Na2SO4→BaSO4+2NaCl $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$ ***BaCl2 +Na2SO4→BaSO4+2NaCl***

  - ***Mass of*** $BaCl_2 + Na_2SO_4$ ***(before reaction) = Mass of*** $BaSO_4 + NaCl$ ***(after reaction).***
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- ***Similarly, when hydrogen burns in oxygen to form water:***

  2H2+O2→2H2O $2H_2 + O_2 \rightarrow 2H_2O$ ***2H2+O2→2H2O***
- ○ ***The total mass of hydrogen + oxygen equals the total mass of water formed.***

Importance:

- ***Proved that matter is indestructible in ordinary chemical changes.***
- ***Helped scientists to understand that chemical reactions only involve rearrangement of atoms, not their destruction.***

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2. Law of Constant Proportion (or Definite Proportion)

- ***Proposed by: Joseph Proust in 1799.***

- ***Statement: "A given compound always contains the same elements combined together in the same fixed proportion by mass, no matter from where it is obtained or how it is prepared."***

Meaning:

- ***A compound has a fixed chemical composition.***

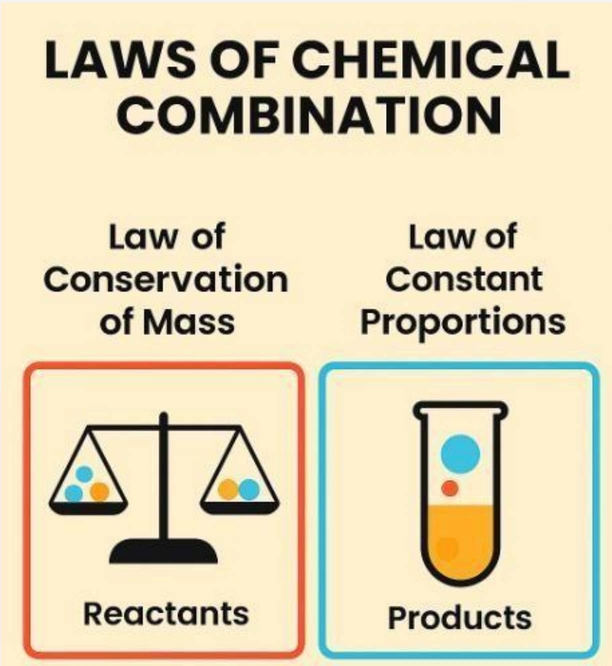
- ***Its identity does not depend on place, source, or method of preparation.***
- ***For example, water will always be H_2O, whether from a river, sea, rain, or made in the lab.***

Examples:

1. ***Water (H_2O):***
    - ***Contains hydrogen and oxygen in the ratio of 1:8 by mass.***
    - ***This ratio is the same for all samples of water.***
2. ***Carbon dioxide (CO_2):***
    - ***Contains carbon and oxygen in the ratio of 3:8 by mass.*** ○ ***A sample from respiration or from burning coal will always show the same proportion.***
3. ***Ammonia (NH_3):***
    - ***Contains nitrogen and hydrogen in the ratio of 14:3 by mass.***

Importance:

- ***Proved that compounds are made up of atoms of elements combined in fixed ratios.***
- ***Supported the idea that matter is not continuous, but made of tiny particles (atoms).***
- ***This law directly led to the development of Dalton's Atomic Theory.***



Summary (Quick Revision):

- ***Law of Conservation of Mass (Lavoisier): Mass is neither created nor destroyed → total mass of reactants = total mass of products.***

- *Law of Constant Proportion (Proust): A compound always contains the same elements in fixed ratio by mass → water is always H_2O, CO_2 is always CO_2.*

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5. Symbols of Elements

To make chemistry simple and universal, every element is represented by a symbol. A symbol is usually the first one or two letters of the element's name, written in a fixed way. For example, Hydrogen → H, Carbon → C, Helium → He*.*

The modern system of symbols was introduced by J.J. Berzelius. The rule is that the first letter is always capital and, if there is a second letter, it is written in lowercase. For instance: Al → Aluminium, Ca → Calcium*.*

Some elements use symbols from their Latin names, like Sodium (Natrium → Na), Potassium (Kalium → K), Iron (Ferrum → Fe), Silver (Argentum → Ag), and Gold (Aurum → Au).

Why are symbols important? They make writing chemical equations easier, remove language barriers, and act as the alphabet of chemistry, since formulas and equations are formed using these symbols.

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6. Atomic Mass

What is Atomic Mass?

- *Every element is made up of tiny particles called atoms.*
- *Different atoms have different weights.*
- *The atomic mass of an element is the average mass of one atom of that element compared to 1/12th the mass of a carbon-12 atom (the standard).*
- *In simple words, atomic mass tells us how heavy an atom is compared to hydrogen or carbon-12.*

Why Do We Need Atomic Mass?

- *Atoms are extremely small and cannot be weighed directly.*
- *Their actual mass is very tiny (≈ 10^{-27} kg).*
- *So, scientists use relative mass instead of actual mass to make calculations simple.*

- ***Atomic mass helps in:***
    - ***Writing formulas of compounds.*** ○ ***Balancing chemical equations.***
    - ***Calculating moles and reacting weights.***

Examples of Atomic Mass

- ***Hydrogen (H): Atomic mass = 1 u (lightest atom).***
- ***Carbon (C): Atomic mass = 12 u.***
- ***Oxygen (O): Atomic mass = 16 u.***
- ***Chlorine (Cl): Atomic mass = 35.5 u (non-whole number because it is an average of isotopes).***

(Here, "u" means unified atomic mass unit.)

Unified Atomic Mass Unit (u)

- ***Earlier, hydrogen was used as the standard, but it was inconvenient.***
- ***Now, scientists use carbon-12 isotope as the reference.***
- ***1 atomic mass unit (1 u) = mass of 1/12th of a carbon-12 atom.***
- ***This makes atomic masses universally accepted and more accurate.***

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7. Molecules

1. Definition

A molecule is the smallest particle of an element or compound that can exist independently and shows all the properties of that substance.

- ***Example: O_2 (oxygen), H_2O (water)***

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2. Types of Molecules

- ***Element molecules: Made of the same atoms*** ○ ***Monoatomic:***

  He, Ne ○ ***Diatomic: H_2, O_2, N_2*** ○ ***Polyatomic: P_4, S_8***
- ***Compound molecules: Made of different atoms in fixed ratios*** ○

  Examples: H_2O, CO_2, NH_3

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3. Key Points

- ***Molecules are held together by chemical bonds.***
- ***They have a definite composition.***
- ***Molecules explain the nature of substances and chemical reactions.***

Takeaway:

Molecules are the building blocks of matter, made of atoms in fixed combinations. They can be simple like O_2 or complex like proteins.

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8. Atomicity

- ***Definition: Number of atoms present in one molecule of an element.***
- ***Examples:***
  - ***Monoatomic: He, Ne, Ar***
  - ***Diatomic: H_2, O_2, N_2***
  - ***Triatomic: O_3***
  - ***Polyatomic: P_4, S_8***

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9. Ions

- ***Charged particles formed by the loss or gain of electrons.***

1. ***Cations (positive ions) – formed by loss of electrons. Example: Na^+, Mg^{2+}.***
2. ***Anions (negative ions) – formed by gain of electrons. Example: Cl^-, SO_4^{2-}.***

10. Chemical Formula

- ***Definition: A symbolic representation of a molecule.***
- ***Formed using:***
    1. ***Symbols of elements.***
    2. ***Valency – combining capacity of an atom.***

Example:
For Aluminium oxide:

Al has valency 3, O has valency 2. Cross-multiplication → Al_2O_3.

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11. Molecular Mass

- ***Sum of atomic masses of all atoms in a molecule.***

  Example: H_2O = (2 × 1) + (1 × 16) = 18 u.

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12. Mole Concept

1. What is a Mole?

- ***A mole is a unit used to measure the amount of a substance.***
- ***1 mole contains 6.022 × 10^23 particles (atoms, molecules, or ions).***
- ***This number is called Avogadro's Number (NA).***
- ***Think of it as a "chemist's dozen": just as 1 dozen = 12 items, 1 mole = 6.022 × 10^23 particles.***

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2. Why Do We Need the Mole Concept?

- ***Atoms and molecules are extremely small, so we cannot weigh them individually.***
- ***The mole lets chemists count atoms and molecules by weighing.***

- ***It links mass of a substance to the number of particles, making chemical calculations possible.***

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3. Molar Mass

- ***Molar mass (M) is the mass of 1 mole of a substance.***
- ***For elements, it is equal to atomic mass in grams.***
  - ***Example: Hydrogen → 1 g/mol, Carbon → 12 g/mol, Oxygen → 16 g/mol***
- ***For compounds, it is the sum of atomic masses of all atoms in the formula. ○ Example: Water (H_2O) → 2×1 + 16 = 18 g/mol ○ Carbon dioxide (CO_2) → 12 + 2×16 = 44 g/mol***

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4. Key Formulae for Calculations

1. ***Number of moles (n): n = Mass of substance ÷ Molar mass***
2. ***Number of particles:***
   Number of particles = n × 6.022 × 10^{23}
3. ***Mass of substance: Mass = n × Molar mass***

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4. ***Examples***

1. Find moles in 18 g of water (H_2O):
- ***Molar mass H_2O = 2×1 + 16 = 18 g/mol • n = 18 ÷ 18 = 1 mole***

2. Number of molecules in 2 moles of O_2:
- ***Number of molecules = 2 × 6.022 × 10^{23} = 1.204 × 10^{24}***

3. Mass of 0.5 mole of CO_2:
- ***Molar mass CO_2 = 12 + 2×16 = 44 g/mol***
- ***Mass = 0.5 × 44 = 22 g 4. Moles in 12 g of carbon (C):***
- ***Atomic mass C = 12 g/mol***
- ***n = 12 ÷ 12 = 1 mole***

5. Number of atoms in 3 moles of sodium (Na):
- ***Number of atoms = 3 × 6.022 × 10^{23} = 1.806 × 10^{24} atoms***

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6. Tips for Solving Numerical Problems

- ***Always check the units: grams, moles, or particles.***
- ***Use atomic or molecular mass from the periodic table.***
- ***Write step-by-step calculations for clarity.***
- ***Remember: 1 mole = 6.022 × 10^{23} particles.***

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13. Numerical Examples

Example 1:

Find the number of molecules in 36 g of water.

- ***Molar mass of H_2O = 18 g/mol.***
- ***Moles = 36 ÷ 18 = 2 moles.***
- ***Molecules = $2 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}$ molecules.***

Example 2:

Calculate mass of 5 moles of CO_2.

- ***Molar mass = 12 + (16 × 2) = 44 g/mol.***
- ***Mass = 5 × 44 = 220 g.***