Chapter -3
Atoms and Molecules

1. Introduction:

- ***Laws of Chemical Combination:*** The chapter begins by establishing the fundamental rules that govern how elements combine to form compounds.
- ***Law of Conservation of Mass*** (by Antoine Lavoisier): Mass can neither be created nor destroyed in a chemical reaction. (or) The total mass of the reactants (substances that react) is always equal to the total mass of the products (substances formed) in a chemical reaction. (or) During a chemical reaction, the sum of the masses of the reactants and products remains unchanged. This is known as the Law of Conservation of Mass.
  Experiment: Burning of magnesium ribbon. The mass of magnesium oxide formed is equal to the mass of magnesium plus the mass of oxygen that combined with it.
- ***Law of Constant Proportions*** (by Joseph Proust): In a chemical substance, the elements are always present in a definite proportion by mass.(or) In a pure chemical compound, elements are always present in a definite proportion by mass. This is known as the Law of Definite Proportions.
  Example: Water (H_2O) is always made up of hydrogen and oxygen in a mass ratio of 1:8, regardless of its source (rain, river, well, etc.).

2. Dalton's Atomic Theory: John Dalton proposed a theory to explain the laws of chemical combination.

Postulates:

1. All matter is made of very tiny particles called atoms, which participate in chemical reactions.
2. Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.
3. Atoms of a given element are identical in mass and chemical properties.
4. Atoms of different elements have different masses and chemical properties.
5. Atoms combine in simple whole-number ratios to form compounds.
6. In a given compound, the relative number and kinds of atoms are constant.

Limitations of Dalton's Atomic Theory:

- Atoms are now known to be divisible, containing subatomic particles (electrons, protons, neutrons).
- Atoms of the same element can have different masses (isotopes).
- It failed to explain the existence of allotropes (e.g., diamond and graphite, both forms of carbon).

3. What is an Atom?

The smallest particle of an element that can exist and may or may not have independent existence.
Atomic radius is measured in nanometres.

$1/10^9$ m = 1 nm

1 m = 10^9 nm

Example: Atomic radius of Hydrogen = 10^{-10} m.

Symbols: Used to denote atoms of elements. Derived from their English or Latin names.

Rules for writing Symbols: First letter is always a capital letter.

* If two letters, the second is a small letter.
* Examples: C (Carbon), Ca (Calcium), Co (Cobalt) - Note: 'Co' is Cobalt, 'CO' is a compound, Carbon Monoxide.

Element	Symbol	Element	Symbol	Element	Symbol
Aluminium	Al	Copper	Cu	Nitrogen	N
Argon	Ar	Fluorine	F	Oxygen	O
Barium	Ba	Gold	Au	Potassium	K
Boron	B	Hydrogen	H	Silicon	Si
Bromine	Br	Iodine	I	Silver	Ag
Calcium	Ca	Iron	Fe	Sodium	Na
Carbon	C	Lead	Pb	Sulphur	S
Chlorine	Cl	Magnesium	Mg	Uranium	U
Cobalt	Co	Neon	Ne	Zinc	Zn

4. Atomic Mass: Since atoms are very light, we use a relative scale.

Definition: Atomic mass is the relative mass of an atom compared to the mass of a carbon-12 atom, which is taken as exactly 12 atomic mass units.

Unit is Atomic Mass Unit (u).

1 u = 1/12th the mass of a carbon-12 atom.

Why Carbon-12?

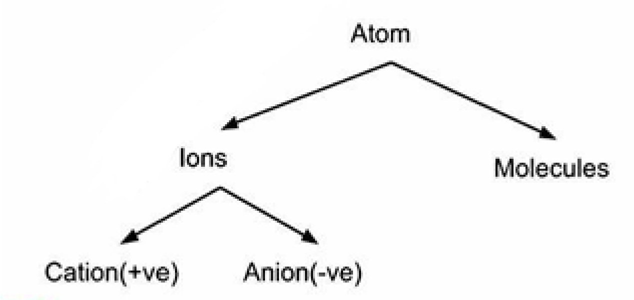
It is chosen as the standard because it is a solid and can be handled easily.

Example: Hydrogen has an atomic mass of ~1 u, oxygen has ~16 u.

Element	Atomic Mass (u)
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Sodium	23
Magnesium	24
Sulphur	32
Chlorine	35.5
Calcium	40

How do atom exist?

Atoms of most elements are not able to exist independently. Atoms form molecules and ions.



5. What is a Molecule?

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

Molecules of Elements:

Formed by atoms of the same element: Can be monoatomic (Noble gases: He, Ne, Ar), diatomic (H_2, N_2, O_2, Cl_2), triatomic (O_3 - Ozone), or tetratomic (P_4 - Phosphorus).

Molecules of Compounds: Formed by atoms of two or more different elements chemically combined in a fixed ratio.

Examples: H_2O (2 H, 1 O), CO_2 (1 C, 2 O), NH_3 (1 N, 3 H).

Atomicity: Number of atoms in a molecule of an element.

Type	Examples	Atomicity
Monoatomic	He, Ne, Ar	1
Diatomic	H_2, O_2, N_2, Cl_2	2
Triatomic	O_3	3
Polyatomic	P_4, S_8	4, 8

6. Ions: Charged particles. They can be positively charged (cations) or negatively charged (anions).

Cations: Formed when an atom loses one or more electrons.

Example: Sodium ion (Na^+), Calcium ion (Ca^{2+}).

Anions: Formed when an atom gains one or more electrons.

Example: Chloride ion (Cl^-), Oxide ion (O^{2-}).

Polyatomic Ions: A group of atoms carrying a charge.

Example: Ammonium (NH_4^+), Hydroxide (OH^-), Carbonate (CO_3^{2-}), Sulphate (SO_4^{2-}).

7. Chemical Formulae: A chemical formula represents the composition of a molecule of a substance.

Rules for Writing Formulae:

Vale-ncy	*Name of ion*	*Symbol*	*Non-metallic element*	*Symbol*	*Polyatomic ions*	*Symbol*
1.	Sodium	Na^+	Hydrogen	H^+	Ammonium	NH_4^+
	Potassium	K^+	Hydride	H^-	Hydroxide	OH^-
	Silver	Ag^+	Chloride	Cl^-	Nitrate	NO_3^-
	Copper (I)*	Cu^+	Bromide	Br^-	Hydrogen carbonate	HCO_3^-
			Iodide	I^-		
2.	Magnesium	Mg^{2+}	Oxide	O^{2-}	Carbonate	CO_3^{2-}
	Calcium	Ca^{2+}	Sulphide	S^{2-}	Sulphite	SO_3^{2-}
	Zinc	Zn^{2+}			Sulphate	SO_4^{2-}
	Iron (II)*	Fe^{2+}				
	Copper (II)*	Cu^{2+}				
3.	Aluminium	Al^{3+}	Nitride	N^{3-}	Phosphate	PO_4^{3-}
	Iron (III)*	Fe^{3+}				

1. The valencies (combining capacity) of the two elements/ions are crossed over to become the subscript of the other.
2. For metals and non-metals, the metal is written first (Ex: NaCl).
3. For compounds with polyatomic ions, the ion is enclosed in brackets if the subscript is more than 1 (Ex: $Ca(OH)_2$).

Formulae of simple compounds:

Examples

1. Formula of hydrogen chloride

Symbol H Cl

Valency 1 1

Formula of the compound would be HCl.

2. Formula of hydrogen sulphide

Symbol H S

Valency 1 2

Formula : H_2S

Some more examples

(a) Formula for aluminium oxide:

Symbol Al O

Charge 3+ 2-

Formula : Al_2O_3

(b) Formula for calcium oxide:

Symbol Ca O

Charge 2+ 2-

Important Formulae to Remember:

Water: H_2O | Ammonia: NH_3 | Carbon Dioxide: CO_2 | Glucose: $C_6H_{12}O_6$

Common Salt: NaCl | Hydrochloric Acid: HCl | Sulphuric Acid: H_2SO_4

8. Molecular Mass: The sum of the atomic masses of all the atoms in a molecule of a substance.

Calculation: It is expressed in atomic mass units (u).

Example 1: (H_2O): (2 × Atomic mass of H) + (1 × Atomic mass of O) = (2 × 1) + (1 × 16) = 18 u

Example 2: (HNO_3): (1 × Atomic mass of H) + (1 × Atomic mass of N) + (3 × Atomic mass of O)
= 1 + 14 + 48 = 63 u

9. Formula of Unit Mass: The formula unit mass of a substance is a sum of the atomic masses of all atoms in a formula unit of a compound.

Example: Sodium chloride as discussed above, has a formula unit NaCl.
Its formula unit mass can be calculated as 1 x 23 + 1 x 35.5 = 58.5 u

10. Key Formulae and Calculations

1. Number of Moles (n) = (Given Mass) / (Molar Mass) or mass / M
2. Number of Particles = Number of Moles × Avogadro's Number = $n \times N_a$ = (mass / Molar Mass) × 6.022×10^{23}`

Example Problem: Calculate the number of moles for 20 g of water (H_2O).

Step 1: Molar mass of H_2O = (2×1) + 16 = 18 g/mol.

Step 2: Number of moles (n) = Given Mass / Molar Mass = 20 g / 18 g/mol ≈ 1.11 moles.