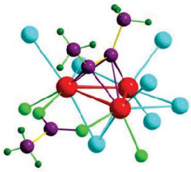


Chapter - 3

Atoms And **Molecules**

CHAPTER AT A GLANCE

Laws of Chemical Combination
- Law of Conservation of Mass
- Law of Constant Proportion
- Law of Multiple Proportion
- Law of Reciprocal Proportion
- Gay Lussac's Law of Combining Volumes

(Law of Multiple Proportion, Law of Reciprocal Proportion, Gay Lussac's Law of Combining Volumes →) Will dealt in higher classes

Dalton's Atomic Theory
- Proves, Law of Conservation of Mass
- Proves, Law of Constant Proportion

Atoms
- Atomic Mass
- Existence of Atoms

Molecules
- Type of Molecules
  - Homoatomic molecules
  - Heteroatomic molecules
- Molecular Mass
- Formula Unit Mass
- Atomicity

Ions

- Simple Ions
- Compound Ions

→ **Mole Concept** → Numericals based upon Mole

Laws of Chemical Combination

The chemical reaction between two or more substances giving rise to products is governed by certain laws. These laws are called 'Laws of Chemical Combination'.

Law of Chemical Combination

- Law of Conservation of Mass
- Law of Constant Proportion
- Law of Multiple Proportion
- Law of Reciprocal Proportion
- Gay Lussac's law of combining volume

(Law of Multiple Proportion, Law of Reciprocal Proportion, Gay Lussac's law of combining volume) → Will study in higher classes

Law of Conservation of Mass

- According to this law, "Mass can neither be created nor destroyed."
- In a chemical reaction, this law can be understood in the following way :

"During a chemical reaction total mass of reactants will be equal to total mass of products."

- For example, $A + B \rightarrow AB$

Reactant Product

Then, $m_A + m_B = m_{AB}$

where, m_A = Mass of A

m_B = Mass of B

m_{AB} = Mass of AB

$2H_2$ (g)	+	O_2 (g)	→	$2H_2O$ (*l*)
$2 \times 2 = 4$ gm		$2 \times 16 = 32$ gm		$2 \times (2 + 16) = 36$ gm

Example : In a reaction 5.3 gm of sodium carbonate reacted with 6 gm of ethanoic acid. The products were 2.2 gm of CO_2, 0.9 gm of H_2O and 8.2 gm of sodium ethanoate. Show that these observation are all in agreement with law of conservation of mass.

Sodium carbonate + Ethanoic acid $\rightarrow$ Sodium ethanoate + CO_2 + H_2O

Solution :

$$\underbrace{\text{Sodium carbonate} + \text{Ethanoic acid}}_{\text{Reactants}} \rightarrow \underbrace{\text{Sodium ethanoate} + CO_2 + H_2O}_{\text{Products}}$$

Now, according to the law of conservation of mass :

Mass of sodium carbonate + Mass of ethanoic acid = Mass of sodium ethanoate + Mass of CO_2 + Mass of H_2O

Putting values of masses from the equation :

$$5.3 \text{ gm} + 6.0 \text{ gm} = 8.2 \text{ gm} + 2.2 \text{ gm} + 0.9 \text{ gm}$$

Or $$11.3 \text{ gm} = 11.3 \text{ gm}$$

Since, LHS = RHS

$\therefore$ Law of conservation of mass is in agreement with the given values in equation.

Law of Constant Proportion

According to this law, "A pure chemical compound always contain the same elements combined together in the same proportion by mass irrespective of the fact from where the sample has been taken or from which procedure has it been produced."

For example :

18 gm of $H_2O \Rightarrow$ 16 gm of oxygen + 2 gm of hydrogen,

i.e., $m_H/m_O = 2/16 = 1/8$

36 gm of $H_2O \Rightarrow$ 32 gm of oxygen + 4 gm of hydrogen,

i.e., $m_H/m_O = 4/32 = 1/8$

09 gm of $H_2O \Rightarrow$ 08 gm of oxygen + 1 gm of hydrogen,

i.e., $m_H/m_O = 1/8$

From the above three cases, differently weighing H_2O samples were taken but the ratio of masses of 'H' to mass of 'O' comes out to be '1/8' is same, proving law of constant proportion.

Likewise, if a sample of 'H_2O' was taken from anywhere *i.e.,* from well, pond, lake or anywhere the ratio of masses of 'H' to 'O' will come out to be same as '1/8'.

Example : Hydrogen and oxygen combine in the ratio 1 : 8 by mass to form water. What mass of oxygen gas would be required to react completely with 3.0 gm of hydrogen gas ?

Solution : $\frac{m_H}{m_O} = \frac{1}{8}$ Given in equation (For H_2O)

But, $m_H = 3.0$ gm (given)

Or $\frac{3}{m_O} = \frac{1}{8}$

Or $m_O = 24$ gm

$\therefore$ Mass of oxygen will be 24 gm.

Or it will be a sample of 27 gm of H_2O where 3 gm of hydrogen is present with 24 gm of oxygen.

Dalton's Atomic Theory

Based upon laws of chemical combination, **Dalton's Atomic Theory** provided an explanation for the **Law of Conservation of Mass** and **Law of Constant Composition.**

Postulates of Dalton's atomic theory are as follows :

- All matter is made up of very tiny particles called 'Atoms'.
- Atom are indivisible particles, which can't be created or destroyed in a chemical reaction. (Proves 'Law of Conservation of Mass')
- Atoms of an element have identical mass and chemical properties.
- Atoms of different elements have different mass and chemical properties.
- Atom combine in the ratio of small whole numbers to form compounds. (Proves 'Law of Constant Proportion')
- The relative number and kinds of atoms are constant in a given compound.

Atom

- According to modern atomic theory, an atom is the smallest particle of an element which takes part in chemical reaction such that during the chemical reaction, the atom maintain its identity, throughout the chemical or physical change.
- Atoms are very small and hence can't be seen even through very powerful microscope.
- Atomic radius of smallest atom in hydrogen is 0.37×10^{-10} m or 0.037 nm.

Such that, $1\ \text{nm} = 10^{-9}\ \text{m}$

IUPAC (International Union of Pure & Applied Chemistry) Symbols of Atoms of Different Elements

Element	Symbol	Element	Symbol
Aluminium	Al	Iodine	I
Argon	Ar	Iron	Fe
Barium	Ba	Lead	Pb
Calcium	Ca	Nitrogen	N
Carbon	C	Oxygen	O
Chlorine	Cl	Potassium	K
Cobalt	Co	Silicon	Si
Copper	Cu	Silver	Ag
Fluorine	F	Sulphur	S
Gold	Au	Zinc	Zn
Hydrogen	H		

Atomic Mass

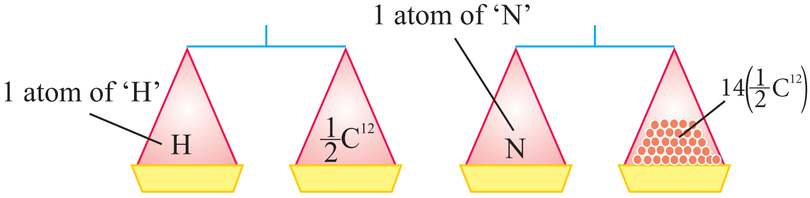
- The mass of an atom of an element is called its atomic mass.
- In 1961, IUPAC have accepted 'atomic mass unit' (u) to express atomic and molecular mass of elements and compounds.

Atomic Mass Unit

The atomic mass unit is defined as the quantity of mass equal to 1/12 of mass of an atom of carbon-12.

$$1\ \text{amu or } u = \frac{1}{12} \times \text{Mass of an atom of } C^{12}$$

$$1\ u = 1.66 \times 10^{-27}\ \text{kg}$$



Atomic Mass of H = 1u Atomic Mass of N = 14 u

Likewise,

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

How do atoms exist ?

- Atoms of most of the elements are very reactive and does not exist in free state.
- Only the atoms of noble gases (such as He, Ne, Ar, Kr, Xe and Rn) are chemically unreactive and can exist in the free state as single atom.
- Atoms of all other elements combine together to form molecules or ions.

Atom (Electrically neutral)

Ion (electrically charged)   Molecules (electrically neutral)

Molecule

- A molecule is a group of two or more atoms which are chemically bonded with each other.
- A molecule is the smallest particle of matter (except element) which is capable of an independent existence and show all properties of that substance.

E.g., 'H_2O' is the smallest particle of water which shows all the properties of water.

- A molecule may have atom of same or different elements, depending upon this, molecule can be categorized into two categories :

Homoatomic molecules (containing atom of same element) and **Heteroatomic molecules** or **compounds** (containing atoms of different elements)

Molecules

- Homoatomic molecules → *E.g.,* O_2, N_2, O_3, S_8, P_4 etc.
- Heteroatomic molecules → *E.g.,* H_2O, CO_2, NaCl, $CaCO_3$ etc.

Atomicity

The number of atoms present in one molecule of an element is called its atomicity.

	Name	Formula	Atomicity	
1.	Argon	Ar	Monoatomic (1)	Noble gasses constitute monoatomic molecules
2.	Helium	He	Monoatomic (1)	
3.	Oxygen	O_2	Diatomic (2)	
4.	Hydrogen	H_2	Diatomic (2)	
5.	Phosphorus	P_4	Tetratomic (4)	
6.	Sulphur	S_8	Polyatomic (8)	

Chemical formulae

It is the symbolic representation of the composition of a compound.

Characteristics of chemical formulae

- The valencies or charges on ion must balance.
- When a compound is formed of metal and non-metal, symbol of metal comes first. *E.g.,* CaO, NaCl, CuO.
- When polyatomic ions are used, the ions are enclosed in brackets before writing the number to show the ratio. *E.g.,* $Ca(OH)_2$, $(NH_4)_2SO_4$

Molecular Mass

It is the sum of atomic masses of all the atoms in a molecule of that substance.

E.g., Molecular mass of H_2O = 2 × Atomic mass of Hydrogen + 1 × Atomic mass of Oxygen

So, Molecular mass of $H_2O = 2 \times 1 + 1 \times 16 = 18\ u$

Formula Unit Mass

It is the sum of atomic mass of ions and atoms present in formula for a compound.

E.g., In NaCl, Na = 23 a.m.u. Cl = 35.5 a.m.u.

So, Formula unit mass $= 1 \times 23 + 1 \times 35.5 = 58.5\ u$

Rules for writing chemical formulae

(i) We first write symbols of elements which form compound.

(ii) Below the symbol of each element, we should write their valency.

(iii) Now cross over the valencies of combining atoms.

(iv) With first atom, we write the valency of second atom (as a subscript).

(v) With second atom, we write the valency of first atom (subscript).

Examples :

(i) Symbol : H S

Valencies : 1 2 H_2S_1 or H_2S (Hydrogen sulphide)

(ii) Symbol : C O

Valencies : 4 2 C_2O_4 or CO_2 (Carbon dioxide)

[Take 2 common and divide the formula by 2]

(iii)For Hydrochloric acid (Hydrogen chloride)

H Cl

1 1 H_1Cl_1 or HCl

(iv)For Carbon tetrachloride

C Cl

4 1 C_1Cl_4 or CCl_4

(v) For Magnesium chloride

Mg Cl

2 1 $MgCl_2$

(vi)For aluminium oxide

Al O

3 2 Al_2O_3

(vii)For Calcium oxide

Ca O

2 2 Ca_2O_2 or CaO

[Take 2 common and divide the formula by 2] (viii)For Sodium nitrate (For ions)

$$\begin{array}{ccc} Na & & NO_3 \\ & \times & \\ 1^{-} & & 1^{-} \end{array} \qquad NaNO_3$$

Ions

An ion may be defined as an atom or group of atoms having positive or negative charge.

Some positively charged ions : Na^{+}, K^{+}, Ca^{2+}, Al^{3+}

Some negatively charged ions : Cl^{-} (chloride ion), S^{2-} (sulphide ion), OH^{-} (hydroxide ion), SO_4^{2-} (sulphate ion)

Ions

Simple ions	Compound ions
Mg^{2+} (Magnesium ion)	NH_4^{+} (Ammonium ion)
Na^{+} (Sodium ion)	CO_3^{2-} (Carbonate ion)
Cl^{-} (Chloride ion)	SO_4^{2-} (Sulphate ion)
Al^{3+} (Aluminium ion)	OH^{-} (Hydroxide ion)

Chemical Formulae of Ionic Compounds (Polyatomic)

(i) Sodium carbonate

$$\begin{array}{ccc} Na & & CO_3 \\ & \times & \\ 1+ & & 2- \end{array} \qquad Na_2CO_3$$

(ii) Aluminium sulphate

$$\begin{array}{ccc} Al & & SO_4 \\ & \times & \\ 3+ & & 2- \end{array} \qquad Al_2(SO_4)_3$$

(iii) Calcium hydroxide

$$\begin{array}{ccc} Ca & & OH \\ & \times & \\ 2+ & & 1 \end{array} \qquad Ca(OH)_2$$

(iv)Ammonium sulphate

$$\begin{array}{ccc} NH_4 & & SO_4 \\ & \times & \\ 1+ & & 2- \end{array} \qquad (NH_4)_2SO_4$$

(v) Magnesium hydroxide

$$\begin{array}{cc} Mg & OH \\ 2+ & 1- \end{array} \qquad Mg(OH)_2$$

Molar Mass

The molar mass of a substance is the mass of 1 mole of that substance.

It is equal to the 6.022×10^{23} atoms of that element/substance.

Example :

(a) Atomic mass of hydrogen (H) is 1 *u*. Its molar mass is 1 g/mol.

(b) Atomic mass of nitrogen is 14 *u*. So, molar mass of nitrogen (N) is 14 g/mol.

(c) Molar mass of S_8 = Mass of S × 8 = 32 × 8 = 256 g/mol

(d) Molar mass of HCl = Mass of H + Mass of Cl

= 1 = 35.5 = 36.5 g/mol

Mole concept

A group of 6.022×10^{23} particles (atoms, molecules or ions) of a substance is called a mole of that substance.

1 mole of atoms = 6.022×10^{23} atoms

1 mole of molecules = 6.022×10^{23} molecules

Example, 1 mole of oxygen = 6.022×10^{23} oxygen atoms

6.022×10^{23} is Avogadro Number (L).

- 1 mole of atoms of an element has a mass equal to gram atomic mass of the element.

Important Formulae

(i) Number of moles (n) = $\dfrac{\text{Given mass}}{\text{Molar mass}} = \dfrac{m}{M}$

(ii) Number of moles (n) = $\dfrac{\text{Given number of particles}}{\text{Avogadro's number}}$

$$n = \frac{N}{N_0}$$

(iii) $\dfrac{m}{M} = \dfrac{N}{N_0}$ $\qquad m = \dfrac{M \times N}{N_0}$

(iv) Percentage of any atom in given compound $= \dfrac{\text{Mass of element} \times 100}{\text{Mass of compound}}$

Example. *Calculate no. of iron atoms in a piece of iron weighing 2.8 gm (At. mass = 54 u).*

Solution : 1 mole of iron = 56 gm (Gram atomic mass of iron)

1 mole of iron element contains 6.022×10^{23} atoms of iron.

So, 56 gm of iron $= 6.022 \times 10^{23}$ atoms

2.8 gm of iron $= \dfrac{6.022 \times 10^{23}}{56} \times 2.8$

$= 3.011 \times 10^{22}$ atoms

Example. *Mass of one molecule of a substance is 5.32×10^{-23} g. What is its molecular mass ?*

Solution : Mass of 1 molecule of substance

$= 5.32 \times 10^{-23}$ g

Mass of 6.022×10^{23} molecules of substance

$= 5.32 \times 10^{-23} \times 6.022 \times 10^{23}$

$= 32$ g

Example. *Calculate the mass of 0.5 mole of N_2 gas.*

Solution : 1 mole of N_2 = Gram molecular mass of N_2

Or 1 mole of N_2 = 28 gm

$\therefore$ 0.5 mole of N_2 gas $= 0.5 \times 28$

= 14 gm of N_2

Example. *Calculate the total number of O_2 molecules present in 8 gm of O_2.*

Solution : Gram molecular mass of O_2

$= 6.022 \times 10^{23}$ O_2 molecules

Or 32 gm of O_2 $= 6.022 \times 10^{23}$ O_2 molecules

Or 8 gm of O_2 $= 6.022 \times 10^{23} \times 8/32$ O_2 molecules

$= 1.51 \times 10^{23}$ O_2 molecules

QUESTIONS

VERY SHORT ANSWER TYPE QUESTIONS

1. Write full form of IUPAC.
2. Name the scientist who gave atomic theory of matter.
3. What are building blocks of matter ?
4. Name two laws of chemical combination.
5. Name the unit in which atomic radius is usually expressed.
6. Define molecular mass.
7. What is formula unit mass ?
8. Name the element used as standard for atomic mass scale.

SHORT ANSWER TYPE QUESTIONS

1. What is atomicity ? Explain with two examples.
2. State law of conservation of mass.
3. State law of constant proportion.
4. Calculate molecular mass of H_2 and NH_3. (At. mass of H = 1 *u*, N = 14 *u*)

LONG ANSWER TYPE QUESTIONS

1. Write postulates of Dalton's atomic theory.
2. What is the difference between molecule of an element and the molecule of a compound ? Give one example of each.

HOTS

1. In what form does oxygen gas occur in nature ?
2. In what form do noble gases occur in nature ?
3. What is the difference between 2H and H_2 ?

NUMERICALS

1. Calculate the gram atomic mass of one atom of oxygen. (Gram atomic mass of oxygen = 16 gm) *[**Ans.** 2.66×10^{-23} gm]*

2. What would be gram atomic mass of 5 moles of chlorine ? *[**Ans.** 177.5 gm]*

3. Calculate the number of moles present in the following :

   (a) 52 gm of He

   (b) 12.044×10^{23} He atoms *[**Ans.** (a) 13 moles, (b) 2 moles]*