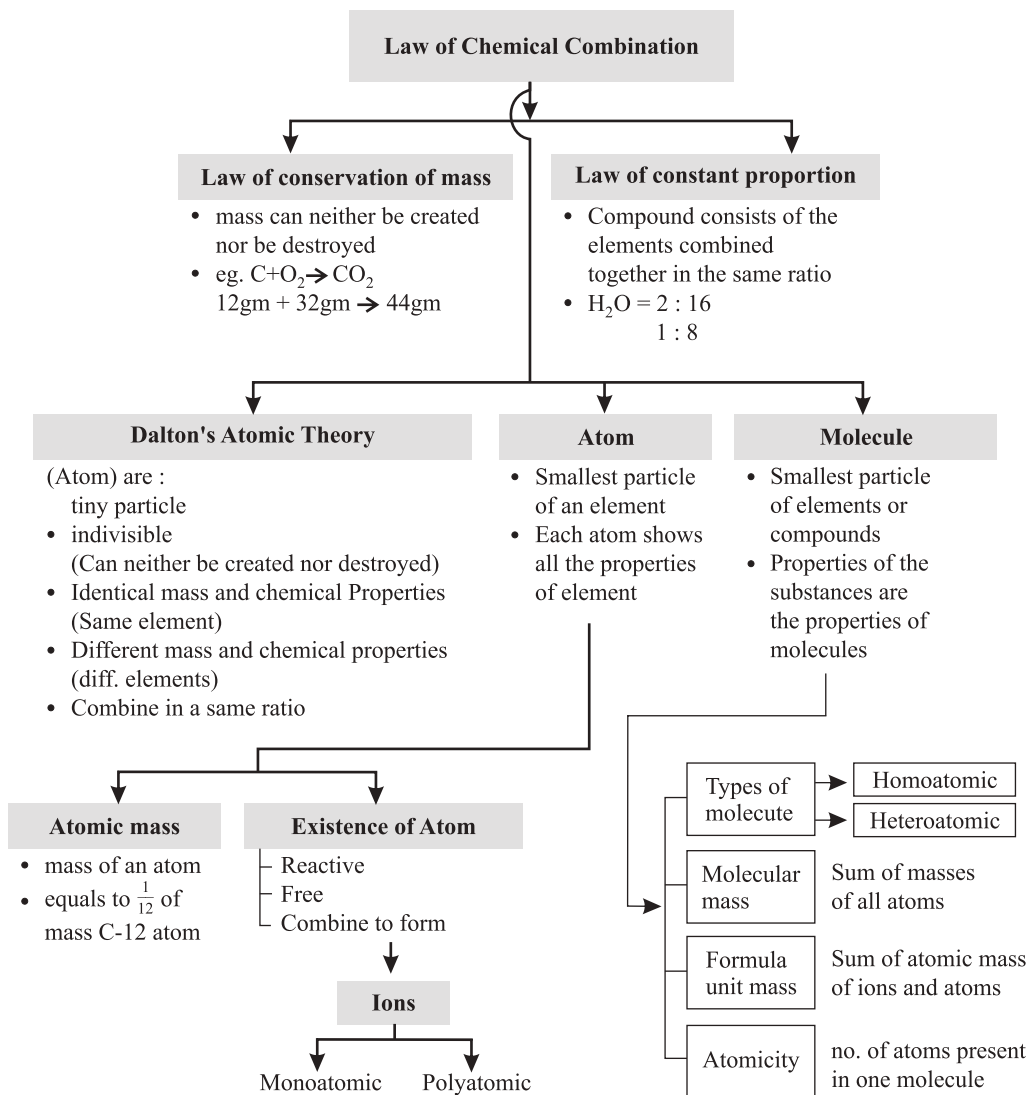




## Chapter - 3

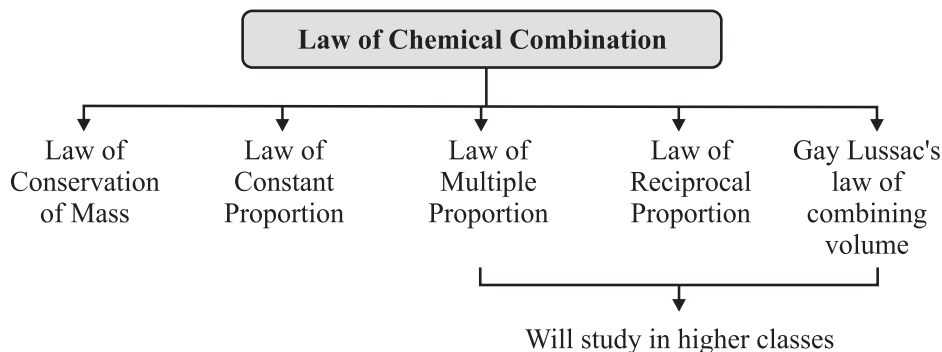
# Atoms And Molecules

### CONCEPT MAPPING



### 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 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,



Reactant                  Product

Then,

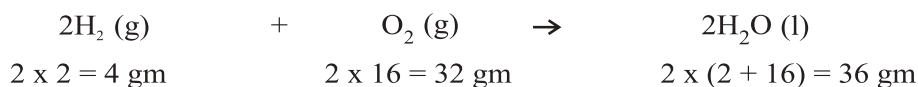
$$m_A + m_B = m_{AB}$$

where,

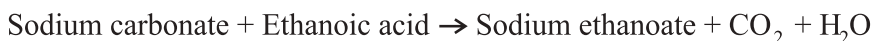
$$m_A = \text{Mass of A}$$

$$m_B = \text{Mass of B}$$

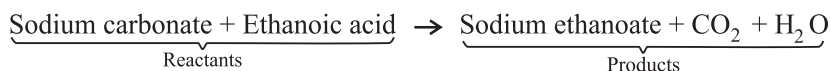
$$m_{AB} = \text{Mass of AB}$$



**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<sub>2</sub>, 0.9 gm of H<sub>2</sub>O and 8.2 gm of sodium ethanoate. Show that these observation are all in agreement with law of conservation of mass.



**Solution :**



Now, according to the law of conservation of mass :

**Mass of sodium carbonate + Mass of ethanoic acid = Mass of sodium ethanoate + Mass of CO<sub>2</sub> + Mass of H<sub>2</sub>O**

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

∴ 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<sub>2</sub>O ⇒ 16 gm of oxygen + 2 gm of hydrogen,

*i.e.*,  $m_{\text{H}}/m_{\text{O}} = 2/16 = 1/8$

36 gm of H<sub>2</sub>O ⇒ 32 gm of oxygen + 4 gm of hydrogen,

*i.e.*,  $m_{\text{H}}/m_{\text{O}} = 4/32 = 1/8$

09 gm pf H<sub>2</sub>O ⇒ 08 gm of oxygen + 1 gm of hydrogen,

*i.e.*,  $m_{\text{H}}/m_{\text{O}} = 1/8$

From the above three cases, differently weighing H, O 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, O' 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<sub>2</sub>O)

But,  $m_H = 3.0 \text{ gm}$  (given)

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

Or  $m_O = 24 \text{ gm}$

∴ Mass of oxygen will be 24 gm.

Or it will be a sample of 27 gm of H<sub>2</sub>O 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  $0.37 \times 10^{-10} \text{ m}$  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

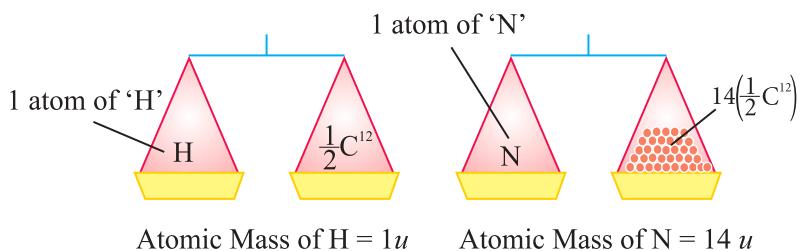
- The mass of an atom of an element is called its atomic mass.
- In 1961, IUPAC have accepted 'atomic mass unit' ( $u$ ) to express atomic and molecular mass of elements and compounds.

### 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.

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

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

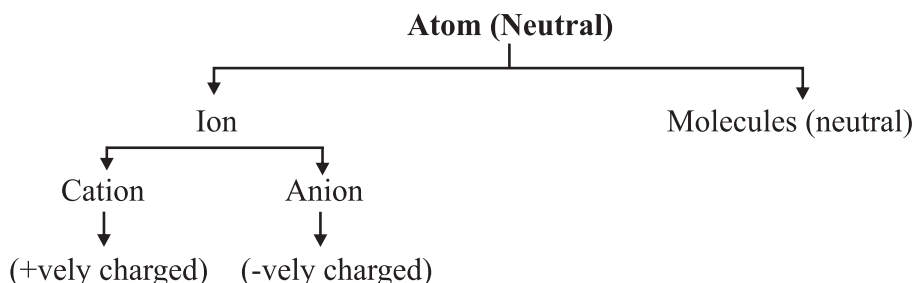


Likewise,

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



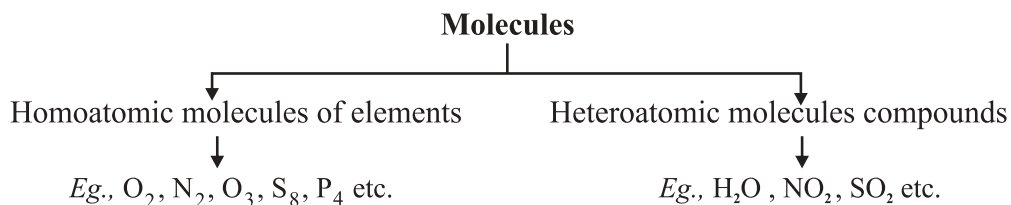
### Molecule

- 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, O' 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)



## 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 <sub>2</sub>	Diatomic (2)	
4. Hydrogen	H <sub>2</sub>	Diatomic (2)	
5. Phosphorus	P <sub>4</sub>	Tetratomic (4)	
6. Sulphur	S <sub>8</sub>	Polyatomic (8)	
7. Ozone	O <sub>3</sub>	Triatomic (3)	

## 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)<sub>2</sub>, (NH<sub>4</sub>)<sub>2</sub> SO<sub>4</sub>

## 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<sub>2</sub>O = 2 x Atomic mass of Hydrogen + 1 x

Atomic mass of Oxygen

So, Molecular mass of H<sub>2</sub>O = 2 x 1 + 1 x 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 x 23 + 1 x 35.5 = 58.5 u

### Rule for writing chemical formulae

Rule-1 (i) We first write symbols of elements which form compound.

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

(iii) Now crisscross 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).

Rule-2 When the valency is one, subscript is not written.

Rule-3 When there are multiple number of polyatomic ion, bracket must be used to separate the Polyatomic ion from subscript.

Examples:



[Take 2 common and divide the formula by 2]

(iii) For Hydrochloric acid (Hydrogen chloride)



(iv) For Carbon tetrachloride



(v) For Magnesium chloride



(vi) For aluminium oxide

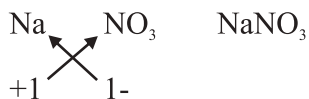


(vii) For Calcium oxide



[Take 2 common and divide the formula by 2]

(viii) For Sodium nitrate (For ions)

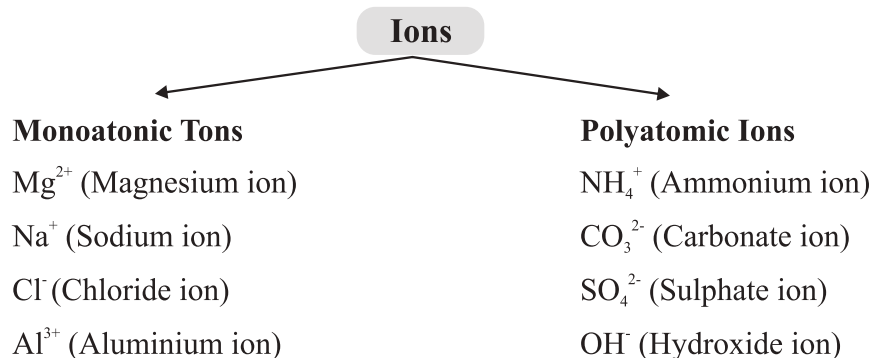


## Ions

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

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

**Some negatively charged ions :**  $\text{Cl}^-$  (chloride ion),  $\text{S}^{2-}$  (sulphide ion),  $\text{OH}^-$  (hydroxide ion),  $\text{SO}_4^{2-}$  (sulphate ion)



## Chemical Formulae of Ionic Compounds (Polyatomic)

(i) Sodium carbonate



(ii) Aluminium sulphate



(iii) Calcium hydroxide



(iv) Ammonium sulphate



(v) Magnesium hydroxide



## Molar Mass

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

It is equal to the  $6.022 \times 10^{23}$  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  $\text{S}_8$  = Mass of S  $\times$  8 = 32  $\times$  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  $6.022 \times 10^{23}$  Particles (atoms, molecules or ions) of a substance is called a mole of that substance.

$$1 \text{ mole of atoms} = 6.022 \times 10^{23} \text{ atoms}$$

$$1 \text{ mole of molecules} = 6.022 \times 10^{23} \text{ molecules}$$

*Example,* 1 mole of oxygen =  $6.022 \times 10^{23}$  oxygen atoms

$6.022 \times 10^{23}$  is Avogadro Number ( $N_0$ ).

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

## Important Formulae

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

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

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

$$(iii) \frac{m}{M} = \frac{N}{N_0} \quad m = \frac{M \times N}{N_0}$$

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

**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  $6.022 \times 10^{23}$  atoms of iron.

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

$$\begin{aligned} 2.8 \text{ gm of iron} &= \frac{6.022 \times 10^{23}}{56} \times 2.8 \\ &= 3.011 \times 10^{22} \text{ atoms} \end{aligned}$$

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

**Solution :** Mass of 1 molecule of substance

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

Mass of  $6.022 \times 10^{23}$  molecules of substance

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

$$= 32 \text{ g}$$

**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 =  $0.5 \times 28$

$$= 14 \text{ gm of } N_2$$

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

**Solution:** Gram molecular mass of  $O_2$

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

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

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