## Molecular Mass

Molecular mass expresses as to how many times a molecule of a substance is heavier than $\frac{1}{12}$ th of the mass of an atom of carbon (carbon-12). Thus, Molecular mass =

$$\frac{\text{Mass of a molecule}}{\frac{1}{12}\text{th mass of a carbon atom (carbon} - 12)}$$

Molecular mass = 2 × vapour density

**Ex.** A molecule of water is 18 times heavier than $\frac{1}{12}$ th of the mass of carbon atom. Therefore, the molecular mass of water is 18 u.

**Calculation of molecular mass from atomic masses**

The molecules are made up of two or more atoms of different elements. Therefore, the molecular mass may be calculated as the sum of the atomic masses of all the atoms in a molecule of that substance.

**Ex.** Ammonia has the formula, $NH_3$. it consists of one atom of N and three atoms of H. The atomic mass of N and H are 14.0 and 1 respectively. Therefore, the molecular mass of $NH_3$ is

Molecular mass of $NH_3$ = At. mass of N + 3 × At. mass of H

= 14 + 3 × 1 = 17 u

**Ex.** Sulphuric acid has the formula $H_2SO_4$. It consists of two H, one S and four O atoms. The atomic masses of H, S and O are 1,32 and 16 respectively. Therefore, the molecular mass of $H_2SO_4$ is

Molecular mass of $H_2SO_4$ =

(2 × at. m. of H) + (1 × at. m. of S)

+ (4 × at. m. of O)

= (2 × 1) + (1 × 32) + (4 × 16) = 98 u

### Gram molecular mass

The molecular mass of a substance expressed in gram is called its gram molecular mass.

**Ex.** Molecular mass of oxygen, $O_2$ = 32 u

So, gram molecular mass of oxygen, $O_2$ = 32 grams.

## Mole Concept

Atoms and molecules are so small in size that they cannot be counted individually. The chemists use the unit mole for counting atoms, molecules or ions. It is represented by n. A mole represents $6.022 \times 10^{23}$ particles.

**Ex.** 1 mole of atoms = $6.022 \times 10^{23}$ atoms.

1 mole of molecules = $6.022 \times 10^{23}$ molecules The number of particles present in 1 mole of any substance is fixed i.e. $6.022 \times 10^{23}$.

This number is called Avogadro constant or Avogadro number.

it is represented by No.

1 mole of atoms = $6.022 \times 10^{23}$ atoms = Gram atomic mass or Molar mass of element

Number of moles = $\dfrac{\text{Mass of element}}{\text{Molar mass}}$

$$n = \frac{m}{M}$$

Number of moles = $\dfrac{\text{Given number of atoms}}{\text{Avogadro number}}$

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

No. of moles = n

Given mass = m

Molar mass = M

Given number of particles = N

Avogadro number of particles = $N_0$

These relations can be interchanged as

Mass of element, $m = n \times M$

or No. of particles of element, $N = n \times N_0$

Similarly,

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

= Gram molecular mass of Molar mass

Number of moles = $\dfrac{\text{Mass of substance}}{\text{Molar mass}}$

$$n = \frac{m}{M}$$

Number of moles

$$= \frac{\text{Given number of molecules}}{\text{Avogadro number}}$$

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

or $m = n \times M$ and $N = n \times N_0$

**Molarity (M): -** Moles of solute is one litre of solution is known as molarity .

$$M = \frac{\text{Number of moles of solute}}{\text{Volume of solution in litre}}$$

**Ex.** An ornament of silver contains 20 g of silver. Calculate the moles of silver present (atomic mass of silver = 180 u)

**Sol.** Moles of silver,

$$n = \frac{m}{M}$$

Mass of silver, m = 20 g,

Molar mass of silver,

M = 108 g

$\therefore$ $n = \frac{20}{108}$ = 0.185 mol.

**Ex.** How many moles of $CO_2$ are present in 51.2 g of it ?

**Sol.** Molecular mass of $CO_2$ = 12 + 2 + 16 = 44 u

Molar mass of $CO_2$ (M) = 44 g

Mass of $CO_2$ (m) = 51.2 g

Moles of $CO_2$,

$$n = \frac{m}{N} = \frac{51.2}{44} = 1.16 \text{ mol.}$$

**Ex.** Calculate the mass of

(i) 0.5 moles of $N_2$ gas

(ii) 0.5 moles of N atoms

**Sol.** (i) 0.5 moles of $N_2$ gas

Mass = Molar mass × Number of moles

$m = M \times n$

M = 28 g, n = 0.5

$\therefore$ m = 28 × 0.5 = 14 g

(ii) Mass = Molar mass × Number of moles

$m = M \times n$

n = 0.5 mole, M = 14 g

m = 14 × 0.5 = 7 g

**Mass percentage of an element from molecular formula :**

The molecular formula of a compound may be defined as the formula which specifies the number of atoms of various elements in the molecule of the compound.

**Ex.** The molecular formula of glucose is $C_6H_{12}O_6$. This indicates that a molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

The mass percentage of each element is then calculated by the following formula : Mass percentage of element X

$$= \frac{\text{Mass of X in one mole}}{\text{Gram molecular mass}} \times 100.$$

**Ex.** Calculate the percentage composition (by mass) of formaldehyde ($CH_2O$).

**Sol.** Molecular mass of formaldehyde,

$CH_2O$ = 12 × 1 + 1 × 2 + 16 × 1 = 30

Mass of one mole of formaldehyde = 30 g

1 Mole of $CH_2O$ contains 1 mole (12 g) of carbon. 2 moles of hydrogen (2 g) and 1 mole of oxygen (16 g)

Percentage of carbon = $\frac{12g}{30g} \times 100 = 40.0\%$

Percentage of hydrogen = $\frac{2g}{30g} \times 100 = 6.7\%$

Percentage of oxygen = $\frac{16g}{30g} \times 100 = 53.3\%$

## Empirical formula

The empirical formula of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound.

**Ex.** The empirical formula of the compound glucose ($C_6H_{12}O_6$), is $CH_2O$ which shows that C, H, and O are present in the simplest ratio of 1 : 2 : 1.

### Rules for writing the empirical formula

The empirical formula is determined by the following steps :

- Divide the percentage of each elements by its atomic mass. This gives the relative number of moles of various elements present in the compound.
- Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.
- Multiply the figures, so obtained by a suitable integer, if necessary, in order to obtain whole number ratio.
- Finally write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the empirical formula of the compound.

**Ex.** A substance, on analysis, gave the following composition : Na = 43.4%, C = 11.3%, O = 45.3%. Calculate its empirical formula

[Atomic masses = Na = 23, C = 12, O = 16]

**Sol.**

| Element | Symbol | % | Atomic mass | Relative number of moles | Simple ratio of moles | Simplest whole no. ratio |
|---|---|---|---|---|---|---|
| Sodium | Na | 43.4 | 23 | $\frac{43.4}{23} = 1.88$ | $\frac{1.88}{0.94} = 2$ | 2 |
| Carbon | C | 11.3 | 12 | $\frac{11.3}{12} = 0.94$ | $\frac{0.94}{0.94} = 1$ | 1 |
| Oxygen | O | 45.3 | 16 | $\frac{45.3}{16} = 2.83$ | $\frac{2.83}{0.94} = 3$ | 3 |

Therefore, the empirical formula is $Na_2CO_3$

**Determination molecular formula :**

Molecular formula = Empirical formula × n

$$n = \frac{\text{Molecular formula}}{\text{Empirical formula}}$$

**Ex.** On heating a sample of $CaCO_3$, volume of $CO_2$ evolved at NTP is 112 cc. Calculate

(i) Weight of $CO_2$ produced

(ii) Weight of $CaCO_3$ taken

(iii) Weight of CaO remaining

**Sol.** (i) Mole of $CO_2$ produced $\frac{112}{22400} = \frac{1}{200}$ mole

mass of $CO_2$ = $\frac{1}{200}$ × 44 = 0.22 gm

(ii) $CaCO_3 \longrightarrow CaO + CO_2$

(1/200 mole)

mole of $CaCO_3$ = $\frac{1}{200}$ mole

mass of $CaCO_3$ = $\frac{1}{200}$ × 100 = 0.5 gm

(iii) mole of CaO produced = $\frac{1}{200}$ mole

mass of CaO = $\frac{1}{200}$ × 56 = 0.28 gm

* Interesting by we can apply

Conversation of mass or wt. of CaO

= wt. of $CaCO_3$ taken – wt. of $CO_2$ produced

= 0.5 – 0.22 = 0.28 gm

**Ex.** Calculate mass of hydrazine $N_2H_4$ obtained when 1.12 litre of $N_2$ taken at NTP reacts with $H_2$ according to $N_2 + 2H_2 \longrightarrow N_2H_4$.

**Sol.** Moles of $N_2$ taken = $\frac{1.12}{22.4} = \frac{1}{20}$

$N_2 + 2H_2 \longrightarrow N_2H_4$

(1/20 mole) (1/20 mole)

mass of $N_2H_4$ = $\frac{1}{20}$ × 32 = 1.6 gm