

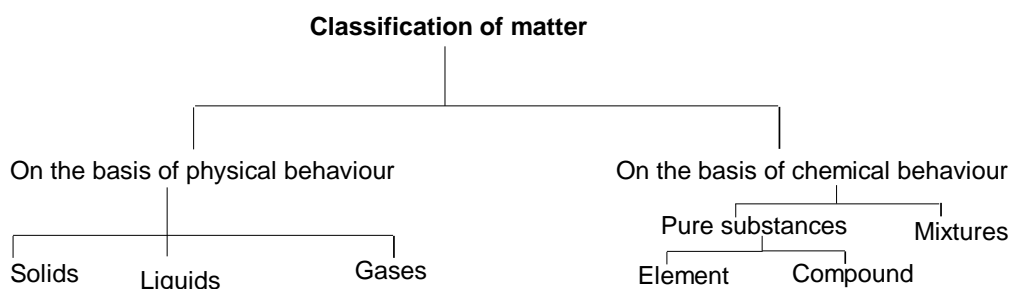
# SOME BASIC CONCEPTS OF CHEMISTRY

## INTRODUCTION

- Anything that occupies space and has mass is called matter.
- Ancient Indian and greek philosopher's believed that the wide variety of object around us are made from combination of five basic elements: Earth, Fire , Water , Air and Sky.
- The Indian philosopher **Kanad** (600 BC) was of the view that matter was composed of very small, indivisible particle called "**parmanus**".
- Ancient Greek philosopher also believed that all matter was composed of tiny building blocks which were hard and indivisible.
- The Greek philosopher Democritus named these building blocks as atoms, meaning indivisible. All these people had their philosophical views about matter, these views were never put to experimental test.
- It was John Dalton who firstly developed a theory on the structure of matter, latter on which was known as **Dalton's atomic theory**.

## DALTON'S ATOMIC THEORY:

- Matter is made up of very small indivisible particle called atoms.
- All the atoms of a element are identical in all respect i.e. mass, shape, size , etc. and atoms of different elements are different in nature.
- Atoms cannot be created or destroyed by any chemical process.



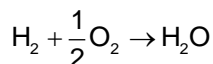
The combination of elements to form compounds is governed by the following five basic laws.

## LAW OF CHEMICAL COMBINATION :

### (a) Law of Conservation of Mass :

**It states that matter can neither be created nor destroyed.**

For any balanced chemical reaction mass of reactant is equal to mass of products. [Exception of law is nuclear reactions where einstein equation is applicable.]



**Example 1:** When 20 g of  $\text{NaHCO}_3$  is heated, 12.62 g of  $\text{Na}_2\text{CO}_3$  and 5.24g of  $\text{CO}_2$  is produced. How many grams of  $\text{H}_2\text{O}$  is produced?

**Solution:** Total mass of  $\text{NaHCO}_3$  heated = 20 gms; Total mass  $\text{Na}_2\text{CO}_3$  produced = 12.62 gms

Total mass of  $\text{CO}_2$  produced = 5.24 gms

$\therefore$  Mass of  $\text{H}_2\text{O}$  produced =  $20 - 12.62 - 5.24 = 2.14$  gms

**(b) The Law of Constant Composition or Definite Proportion (Proust in 1799) :** This law states that " All pure samples of the same chemical compound contain the same elements combined in the same proportion by mass irrespective of the method of preparation".

Example : Different samples of carbon dioxide contain carbon and oxygen in the ratio of 3 : 8 by mass.

Similarly in water ratio of weight of hydrogen to oxygen is 1 : 8.

**Example 2:** When 50 g of ammonia is heated it gives 41.18 g of Nitrogen. When 10 g. of Nitrogen is combined with required amount of hydrogen it produces 12.14g ammonia. Show that the given data follows the law of constant compositions.

**Solution:** If 50g of Ammonia gives 41.18g of Nitrogen, then the percentage of Nitrogen in ammonia

$$\text{is } \frac{41.18}{50} \times 100 = 82.36\%.$$

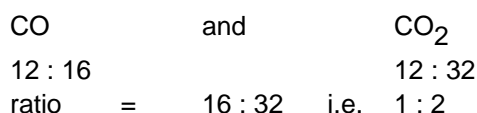
If 10g of Nitrogen gives 12.14 g of Ammonia then percentage of Nitrogen in ammonia is

$$\frac{10}{12.14} \times 100 = 82.37\%.$$

**(c) The Law of Multiple Proportion (Dalton)**

This law states that :when two elements A and B combine together to form more than one compound, then several, masses of A which separately combine with a fixed mass of B, are in a simple ratio.

**Example :**



**(d) The Law of Reciprocal Proportions (Richer in 1792 - 94)**

This law states that “when two elements combine separately with third element and form different types of molecules their combining ratio is directly reciprocated if they combine directly.

**Example :** C combines with O to form CO<sub>2</sub> and with H to form CH<sub>4</sub>. In CO<sub>2</sub> 12 g of C reacts with 32g of O, whereas in CH<sub>4</sub> 12 g of reacts with 4g of H. Therefore when O combines with H, they should combine in the ratio of 32 : 4 (i.e. 8 : 1) or in simple multiple of it. The same is found to be true in H<sub>2</sub>O molecules. The ratio of weight of H and O in H<sub>2</sub>O is 1 : 8.

**(e) The Law of Gaseous Volume (Gay Lussac in 1808)**

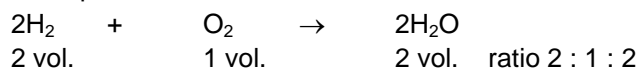
This law states that “when gas combine, they do so in volume which bear a simple ratio to each other and also to the product formed provided all gases are measured under similar conditions. Or in other words volume of reacting gases and product gases have a simple numerical ratio to one another.

Example



**(f) The Avogadro Law**

This law states that “equal volume of all gaseous under similar conditions of temperature and pressure contain equal number of molecules”.



It provides a relationship between vapour density and molecular mass of substances.

$$2 \times \text{vapour density (VD)} = \text{molecular mass of gas.}$$

**ATOMIC AND MOLECULAR MASSES**

**Atomic Mass :**

$$\text{Atomic mass} = \frac{\text{Average mass of an atom}}{1/12 \times \text{Mass of an atom of C}^{12}}$$

**Gram Atomic Mass :**

The atomic mass of an element expressed in grams is called gram atomic mass of that element.

It is also defined as mass of 6.022 x 10<sup>23</sup> atoms.

It is also defined as the mass of one mole atoms.

It is also defined as the mass of 1 gram atom of the element.

**Atomic mass unit (amu) :**

Atomic mass Unit (amu) =  $\frac{1}{12}$  the mass of a carbon - 12 atom

1amu =  $1.660539 \times 10^{-24}$  gm.

Mass of one  $C_6^{12}$  atom = 12 a.m.u

1 a.m.u = mass of one  $C_6^{12}$  atoms =  $\frac{12}{N_0 \times 12} = \frac{1}{N_0}$

Nowadays amu has been replaced by 'u' which is known as unified mass.

**Molecular mass :**

Molecular mass of a molecule, of an element or a compound may be defined as a number which indicates how many times heavier is a molecule of that element or compound as compared with  $\frac{1}{12}$  of the mass of an atom of carbon-12. Molecular mass is a ratio and hence has no units. It is expressed in a.m.u.

Molecular mass =  $\frac{\text{Mass of one molecule of the substance}}{1/12 \times \text{Mass of one atom of C-12}}$

Actual mass of one molecule = Mol. mass  $\times 1.66 \times 10^{-24}$  gm.

**Gram Molecular Mass :**

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

It is also defined as mass of  $6.022 \times 10^{23}$  molecules .

It is also defined as the mass of 1 mole molecules.

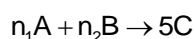
It is also defined as the mass of 1 gram molecule.

**Average atomic mass and molecular mass**

$\bar{A}$  (Average atomic mass) =  $\frac{\sum A_i X_i}{\sum X_{\text{total}}}$

$\bar{M}$  (Average molecular mass) =  $\frac{\sum M_i X_i}{\sum X_{\text{total}}}$

Where  $A_1, A_2, A_3, \dots$  are atomic mass of species 1, 2, 3, .... etc. with % ratio as  $X_1, X_2, X_3, \dots$  etc. Similar terms are for molecular masses.

**Equivalent concept :**

$$\frac{w}{E} = \text{no. of equivalent} = \frac{w}{m} \times n = n \times \text{no. of moles}$$

no. of equivalent of A = no. of equivalent of B = no. of equivalent of C

**Equivalent mass (E) :**

Number of moles of a species =  $\frac{\text{weight (grams)}}{\text{Atomic or molecular mass (g / mole)}} = \frac{w}{M}$

Number of moles of a gas =  $\frac{\text{Volume occupied by gas at NTP}}{\text{Volume occupied by 1 mole of the gas at NTP}}$

$\therefore$  Number of equivalents of solute = n  $\times$  number of moles of solute

$$\text{Also, } N = \frac{w}{M_1/n} \times \frac{1}{v \text{ (in litre)}} = \frac{w}{M_1} \times \frac{1}{V \text{ (in litre)}} \times n; \quad N = M \times n$$

∴ **Normality of solution = n × molarity of solution**

### Concept of minimum molecular mass :

$$\text{Minimum molecular mass} = \frac{A \times 100}{\text{percentage of element}}; \quad A = \text{Atomic mass}$$

### Some Basic units :

% wt or w/w → gm quantity of solute present 100gm of solution

% by v or v/v → volume of solute present in 100 ml of solution

w/v → gm quantity of solute present in 1000ml of solution its unit is gm/lt.

**Example 3 :** The molecular mass of H<sub>2</sub>SO<sub>4</sub> is 98 amu. Calculate the number of moles of each element in 294 g of H<sub>2</sub>SO<sub>4</sub>.

**Solution :** Gram molecular mass of H<sub>2</sub>SO<sub>4</sub> = 98 gm

$$\text{Moles of H}_2\text{SO}_4 = \frac{294}{98} = 3 \text{ moles}$$

H <sub>2</sub> SO <sub>4</sub>	H	S	O
one molecule	2 atoms	one atom	4 atoms
1 × N <sub>A</sub>	2 × N <sub>A</sub> atoms	1 × N <sub>A</sub> atoms	4 × N <sub>A</sub> atoms
∴ one mole	2 mole	one mole	4 mole
∴ 3 mole	6 mole	3 mole	12 mole

### MOLE CONCEPT

- Mole means heap or collection of things. It is basic measuring unit of chemical substance.
- One mole of any chemical substance contains fixed no. of entity i.e. particles and known as Avogadro number (N<sub>o</sub> or N<sub>A</sub>) equal to 6.023 × 10<sup>23</sup>.
- For element :  ${}_7\text{N}^{14}$   
1 mol of N = 14 gm of N = 1 gm atom = N<sub>o</sub> atoms
- For molecule/compound N<sub>2</sub>, NO<sub>2</sub> etc.  
1 mol of N<sub>2</sub> = 28gm of N<sub>2</sub> = 1 gm molecule of N<sub>2</sub> = N<sub>o</sub> molecule = 2 N<sub>o</sub> atoms
- For ions N<sup>3-</sup>, Cl<sup>-</sup>  
1 mol of N<sup>3-</sup> = 14 gm of N<sup>3-</sup> = 1 gm ion of N = N<sub>o</sub> ions

$\text{no. of moles} = n = \frac{w \rightarrow \text{gm quantity}}{m \rightarrow \text{Atomic mass/molecular mass}}$
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- 1 mol of any gas contains fixed volume i.e: 22.4 lt at NTP (0°C & 1 bar)

$$\text{Number of moles of a species} = \frac{\text{weight (grams)}}{\text{Atomic or molecular mass (g / mole)}} = \frac{w}{M}$$

$$\text{Number of moles of a gas} = \frac{\text{Volume occupied by gas at NTP}}{\text{Volume occupied by 1 mole of the gas at NTP}}$$

**Avogadro's hypothesis :** Equal volume of the gases have equal number of molecules ( not atoms) at same temperature and pressure condition.

**S.T.P. :** (Standard Temperature and Pressure)

At S.T. P. condition

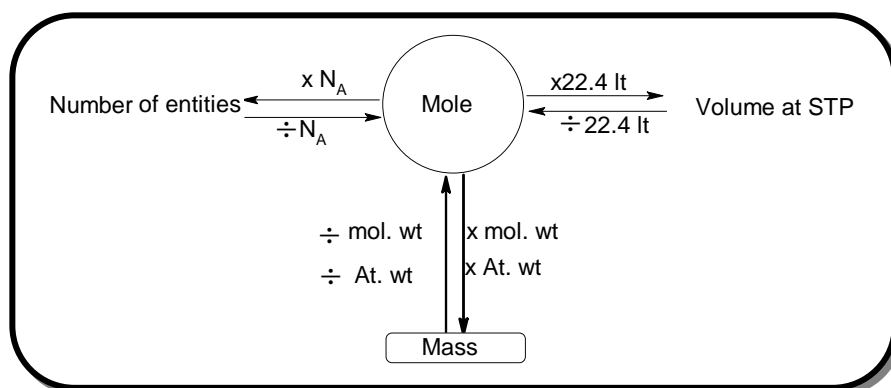
Temperature = 0°C or 273 K

pressure = 1 atm = 760 mm of Hg

and volume of one mole of gas at STP is found to be experimentally equal to 22.4 litres which is known as molar volume.

Avogadro number =  $6.023 \times 10^{23}$

### T-map: Interconversion of mole-volume, mass and number of particles



### Empirical and Molecular Formula :

(i) Dividing % by atomic mass gives molar ratio from which empirical formula is obtained.

(ii) 
$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}}$$

Molecular formula = (Empirical formula)<sub>n</sub>

(iii) Molecular mass = mass of 22.4 l of gas or vapour at S.T.P

**Example 4 :** A substance, on analysis, gave the following percentage composition: Na = 43.4%, C = 11.3%, O = 45.3%. Calculate its empirical formula. [Na = 23, C = 12, O = 16]

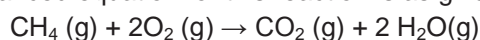
**Solution:**

Element	SYMBOL	% age	Atomic Mass	Relative number moles	of	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  $\text{Na}_2\text{CO}_3$ .

### STOICHIOMETRY

The word 'stoichiometry' is derived from two Greek words - stoicheion (meaning element) and metron (meaning measure). Stoichiometry, thus, deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction. Before understanding how to calculate the amounts of reactants required or those produced in a chemical reaction, let us study what information is available from the balanced chemical equation of a given reaction. Let us consider the combustion of methane. A balanced equation for this reaction is as given below:



The above balanced reaction gives the following information:

- For every 1 mole of  $\text{CH}_4$ , 2 mole of  $\text{O}_2$  will be required to produce 1 mole of  $\text{CO}_2$  and 2 moles of  $\text{H}_2\text{O}$ . this signifies **Mole – Mole relation**
- For every 16 gms of  $\text{CH}_4$ , 64 gms of  $\text{O}_2$  will be required to produce 44gms of  $\text{CO}_2$  and 36 gms of  $\text{H}_2\text{O}$  this signifies **Mass – Mass relation**
- Ratio of moles of  $\text{CO}_2 : \text{H}_2\text{O}$  at any time = 1 : 2

- d) There will be no change in total mass of all reactants and products at any time for any chemical reaction.
- e) For the above reaction only, there will be no change in total number of moles of all reactants and products.

In order to solve the problems based on chemical calculations the following steps, in general, are quite helpful.

- (i) Write the balanced chemical equation.
- (ii) Write the atomic mass/molecular mass/moles/molar volumes of the species involved in calculations.
- (iii) Calculate the result by applying unitary method.

### INTERPRETATION OF BALANCED CHEMICAL EQUATIONS

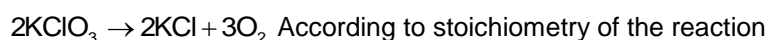
Once we get a balanced chemical equation then we can interpret a chemical equation by following ways

- Mass – mass analysis
- Mass – volume analysis
- Volume – volume analysis

Now you can understand the above analysis by following example

- **Mass – mass analysis**

Consider the reaction



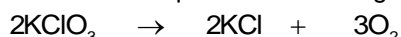
Mass – mass ratio :  $2 \times 122.5$  :  $2 \times 74.5$  :  $3 \times 32$

Or 
$$\frac{\text{Mass of KClO}_3}{\text{Mass of KCl}} = \frac{2 \times 122.5}{2 \times 74.5}$$

$$\frac{\text{Mass of KClO}_3}{\text{Mass of O}_2} = \frac{2 \times 122.5}{3 \times 32}$$

**Example 5 :** 367.5 gram  $\text{KClO}_3$  (  $M = 122.5$ ) is heated. How many gram KCl and oxygen is produced.

**Solution :** Balanced chemical equation for heating of  $\text{KClO}_3$  is



Mass – mass ratio:  $2 \times 122.5$  gm :  $2 \times 74.5$  gm :  $3 \times 32$  gm

$$\frac{\text{Mass of KClO}_3}{\text{Mass of KCl}} = \frac{2 \times 122.5}{2 \times 74.5} \Rightarrow \frac{367.5}{W} = \frac{122.5}{74.5}$$

$$W_{\text{KCl}} = 3 \times 74.5 = 223.5 \text{ gm}$$

$$\frac{\text{Mass of KClO}_3}{\text{Mass of O}_2} = \frac{2 \times 122.5}{3 \times 32} \Rightarrow \frac{367.5}{W} = \frac{2 \times 122.5}{3 \times 32}$$

$$W_{\text{oxygen}} = 144 \text{ gm}$$

- **Mass – volume analysis :**

Now again consider decomposition of  $\text{KClO}_3$



mass volume ratio  $2 \times 122.5$  gm :  $2 \times 74.5$  gm :  $3 \times 22.4$  lt at S.T.P.

we can use two relation for volume of oxygen.

$$\frac{\text{Mass of KClO}_3}{\text{volume of O}_2 \text{ at STP}} = \frac{2 \times 122.5}{3 \times 22.4 \text{ lt}} \dots(\text{i})$$

and 
$$\frac{\text{Mass of KCl}}{\text{volume of O}_2 \text{ at STP}} = \frac{2 \times 74.5}{3 \times 22.4 \text{ lt}} \dots(\text{ii})$$

**Example 6 :** Calculate the volume of  $\text{O}_2$  and volume of air needed for combustion of 1 kg carbon at STP.

**Solution:** 
$$\text{C} + \text{O}_2 \longrightarrow \text{CO}_2$$

$\therefore$  12 g C requires  $O_2 = 22.4$  litre of  $O_2 = 1$  mole of  $O_2 = 32$  g of  $O_2$

$\therefore$  1000 g C requires  $O_2 = \frac{22.4 \times 1000}{12}$  litre  
 $= 1866.67$  litre  $O_2$

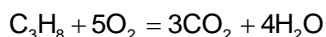
$\therefore V_{\text{air}} = 5 \times V_{O_2} = 5 \times 1866.67 = 9333.35$  litre

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- **Volume – Volume Relationship** : It relates the volume of gaseous species ( reactants or product ) with the volume of another gaseous species ( reactant or product ) involved in a chemical reaction.
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**Example 7:** What volume of oxygen gas at NTP is necessary for complete combustion of 20 litre of propane measured at  $0^\circ\text{C}$  and 760 mm pressure.

**Solution :** The balanced equation is



1vol    5vol

1litre   5litre

1 litre of propane requires = 5 litre of oxygen

20 litre of propane will require =  $5 \times 20 = 100$  litre of oxygen at 760 mm pressure and  $0^\circ\text{C}$ .

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## CONCENTRATION TERMS

**Molarity (M)** : Number of moles present in one lt of solution (mol/lt)

$$\text{Molarity} = \frac{W}{m \times V(\text{ml})} \times 1000$$

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**Example 8:** 149 gm of potassium chloride (KCl) is dissolved in 10 Lt of an aqueous solution. Determine the molarity of the solution (K = 39, Cl = 35.5)

**Solution :** Molecular mass of KCl =  $39 + 35.5 = 74.5$  gm

$\therefore$  Moles of KCl =  $\frac{149\text{gm}}{74.5\text{gm}} = 2$

$\therefore$  Molarity of the solution =  $\frac{2}{10} = 0.2\text{M}$

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**Normality (N)** : Number of equivalent (w/E) present in one lt of solution

$$\text{Normality} = \frac{W}{E \times V(\text{ml})} \times 1000$$

**Molality (m)** : Number of moles of solute present in 1000 gm of solvent known as molality

$$\text{Molality} = \frac{W}{m \times W \rightarrow \text{gm quantity of solvent}} \times 1000$$

**Molarity(M) and Molality(m) for Pure Substances:**

1. **Water** : Let the sample of water has 1000 ml

Mass of water = 1000 gm [density of water = 1gm/mL.]

Moles of water =  $\frac{1000}{18}$  mol

$\therefore$  Molarity =  $\frac{\left(\frac{1000}{18}\right)}{1} = 55.55\text{M}$  and molality =  $\frac{\left(\frac{1000}{18}\right)\text{mol}}{1\text{ kg}} = 55.55\text{ m}$

2. **Pure ethanol** : d gm/ml (density of ethanol)

( $C_2H_5OH$ ) let volume of ethanol taken be 1000 ml.

$\therefore$  wt of ethanol in 1000 ml =  $1000 \times d$  gm

$$\text{Mol of ethanol} = \frac{1000d}{46} \quad \therefore \text{Molarity} = \frac{1000d}{46}$$

$$\text{\& molality of ethanol} = \frac{\left(\frac{1000d}{46}\right)\text{mol}}{\left(\frac{1000d}{1000}\right)\text{kg}} = \frac{1000}{46}$$

- **Parts per million (ppm)**  $\rightarrow$  Amount of solute ( in g ) with  $10^6$  g solvent
- **Parts per billion (ppb)**  $\rightarrow$  amount of solute ( in g ) with  $10^9$  g solvent

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**Example 9 :** 255 gm of an aqueous solution contains 5 gm of urea. What is the concentration of the solution in terms of molality. (Mol. wt. of urea = 60)

**Solution :** Mass of urea = 5 gm

Molecular mass of urea = 60

$$\text{Number of moles of urea} = \frac{5}{60} = 0.083$$

Mass of solvent =  $(255 - 5) = 250$  gm

$$\begin{aligned} \therefore \text{Molality of the solution} &= \frac{\text{Number of moles of solute}}{\text{Mass of solvent in gram}} \times 1000 \\ &= \frac{0.083}{250} \times 1000 = 0.332 \end{aligned}$$

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### Mole Fraction (X)

The ratio of number of moles of the solute or solvent present in the solution and the total number of moles present in the solution is known as the mole fraction of substance concerned.

Let number of moles of solute in solution = n

Number of moles of solvent in solution = N

$$\therefore \text{Mole fraction of solute } (X_1) = \frac{n}{n+N}$$

$$\therefore \text{Mole fraction of solvent } (X_2) = \frac{N}{n+N}$$

$$\text{Also } X_1 + X_2 = 1$$

- Mole fraction is a pure number. It will remain independent of temperature changes.

### Formality (F) :

Number of gram formula weight of a solute dissolve per litre of the solution.

$$= \frac{\text{mass of solute (g)}}{\text{formula mass of solute}} \times \frac{1}{\text{Volume of solution (L)}}$$

### PERCENTAGE CONCENTRATION

The concentration of a solution may also be expressed in terms of percentage in the following way.

- **% weight by weight (w/w):** it is given as mass of solute present in per 100 gm of solution.

$$\text{i.e. } \% \text{ w/w} = \frac{\text{mass of solute in gm}}{\text{mass of solution in gm}} \times 100$$

- **% weight by volume (w/v) :** It is given as mass of solute present in per 100 ml of solution



i.e.  $\% w/v = \frac{\text{mass of solute in gm}}{\text{volume of solution in ml}} \times 100$

- **% volume by volume (V/V) : It is given as volume of solute present in per 100 ml solution.**

i.e.  $\% V/V = \frac{\text{Volume of solute in ml}}{\text{Volume of solution in ml}} \times 100$

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**Example 10 : 0.5 g of a substance is dissolved in 25 g of a solvent. Calculate the percentage amount of the substance in the solution.**

**Solution** : Mass of substance = 0.5 g  
Mass of solvent = 25 g

$\therefore$  Percentage of the substance (w/w) =  $\frac{0.5}{0.5 + 25} \times 100 = 1.96$

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