SOME BASIC CONCEPTS OF CHEMISTRY
LAWS OF CHEMICAL COMBINATIONS

❖ INTRODUCTION

There are a large number of objects around us which we can see and feel.

Anything that Occupies Space and has Mass is called Matter.

It was John Dalton who firstly developed a theory on the structure of matter, later on which is known as Dalton's atomic theory.

1. DALTON'S ATOMIC THEORY

1. Matter is made up of very small undivisible particle called atoms.
2. All the atoms of a given element is identical in all respect i.e. mass, shape, size, etc.
3. Atoms cannot be created nor destroyed by any chemical process.
4. Atoms of different elements are different in nature.

2. THE LAW OF CHEMICAL COMBINATION

In order to understand the composition of various compounds, it is necessary to have a theory which accounts for both qualitative and quantitative observations during chemical changes.

- Atoine Lavoisier, John Dalton and other scientists formulate certain law concerning the composition of matter and chemical reactions. These laws are known as the law of chemical combination.

3. THE LAW OF CONSERVATION OF MASS

It is given by Lavoisier.

It is also known as the law of indestructibility of matter. According to this law, in all chemical changes, the total mass of a system remains constant or in a chemical change, mass is neither created nor destroyed.

In a chemical change total mass remains conserved. i.e., mass before reaction is always equal to mass after reaction. All chemical reactions follow this law. Thus, this law is the basis of all quantitative work in chemistry

Thus, in this chemical change,

Total masses of reactants = 'total masses of products

This relationship holds good when reactants are completely converted into products.

In case, the reacting materials are not completely consumed, the relationship will be

Total masses of reactants = Total masses of products + Masses of unreacted reactants

In a chemical change total mass remains conserved i.e., mass before the reaction is always equal to mass after the reaction.

$$H_2 \quad + \quad 1/2\ O_2 \quad \rightarrow \quad H_2O\ (\lambda)$$

(g) (g)

1 mole1/2 mole 1 mole

mass before the reaction = 1 × 2 + 1/2 × 32 = 18 gm

mass after the reaction = 1 × 18 = 18 gm

Ex. A 15.9g sample of sodium carbonate is added to a solution of acetic acid weighing 20.0g. The two substances react, releasing carbon dioxide gas to the atmosphere. After reaction, the contents of the reaction vessel weigh 29.3g. What is the mass of carbon dioxide given off during the reaction?

Sol. The total mass of reactants taken = 15.9 + 20.0 = 35.9 gm. From the conservation of mass, the final mass of the contents of the vessel should also be 35.9 gm. But it is only 29.3 gm. The difference is due to the mass of released carbon dioxide gas.

Hence, the mass of carbon dioxide gas released = 35.9 – 29.3 = 6.6 gm

Ex.

$$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \longrightarrow H_2O\ (\lambda)$$

Before reaction initially $\frac{1}{2}$ 1 mole mole

After the reaction 0 0 1 mole

mass before reaction = mass of 1 $\frac{1}{2}$ mole H_2 (g) + mole O_2 (g)

= 2 + 16 = 18 g

mass after reaction = mass of 1 mole water = 18 g

4. LAW OF CONSTANT OR DEFINITE PROPORTION

This law was presented by Proust in 1799 and may be stated as follows:

A chemical compound always contains the same element combined together in fixed proportion by mass, i.e., a chemical compound has a f IX ed composition and it does not depend on the method of its preparation or the source from which it has been obtained.

All chemical compounds are found to have constant composition irrespective of their method of preparation or sources.

Ex. In water (H_2O), Hydrogen and Oxygen combine in 2 : 1 molar ratio, the ratio remains constant whether it is tap water, river water or sea water or produced by any chemical reaction.

Ex. 1.80 g of a certain metal burnt in oxygen gave 3.0 g of its oxide. 1.50 g of the same metal heated in steam gave 2.50 g of its oxide. Show that these results illustrate the law of constant proportion.

Sol. In the first sample of the oxide,

wt. of metal = 1.80 g,

wt. of oxygen = (3.0 – 1.80) g = 1.2 g

$$\therefore \frac{\text{wt.of metal}}{\text{wt.of oxygen}} = \frac{1.80\ \text{g}}{1.2\ \text{g}} = 1.5$$

In the second sample of the oxide,

wt. of metal = 1.50 g,

wt. of oxygen = (2.50 – 1.50) g = 1 g

$$\therefore \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.50\ \text{g}}{1\ \text{g}} = 1.5$$

Thus, in both samples of the oxide the proportions of the weights of the metal and oxygen are fixed. Hence, the results follow the law of constant proportion.

5. THE LAW OF MULTIPLE PROPORTION

This law was put forward by Dalton in 1808. According to this law, if two elements combine to form more than one compound, then the different masses of one element which combine with a f IX ed mass of the other element, bear a simple ratio to one another.

..Hydrogen and oxygen combine to form two compounds H^2O (water) and H^20_2 (hydrogen peroxide).

When one element combines with the other element to form two or more different compounds, the mass of one element, which combines with a constant mass of the other, bear a simple ratio to one another.

Note: Simple ratio here means the ratio between small natural numbers, such as 1 : 1, 1 : 2, 1 : 3, Later on this simple ratio becomes the valency and then oxidation state of the element.

Ex: Carbon and Oxygen when combine, can form two oxides viz CO (carbon monoxide), CO_2 (Carbon dioxides)

In CO, 12 g carbon combined with 16 g of oxygen.

In CO_2, 12 g carbon combined with 32 g of oxygen.

Thus, we can see the mass of oxygen which combine with a constant mass of carbon (12 g) bear simple ratio of 16 : 32 or 1 : 2.

Ex. Carbon is found to form two oxides, which contain 42.9% and 27.3% of carbon respectively. Show that these figures illustrate the law of multiple proportions.

Sol. Step-1

To calculate the percentage composition of carbon and oxygen in each of the two oxides.

	First oxide	Second oxide	
Carbon	42.9 %	27.3 %	(Given)
Oxygen	57.1%	72.7 %	(By difference)

Step-2

To calculate the masses of carbon which combine with a fixed mass i.e., one part by mass of oxygen in each of the two oxides.

In the first oxide, 57.1 parts by mass of oxygen combine with carbon = 42.9 parts.

$\therefore$ 1 part by mass of oxygen will combine with carbon $= \frac{42.9}{57.1} = 0.751$.

In the second oxide. 72.7 parts by mass of oxygen combine with carbon = 27.3 parts.

$\therefore$ 1 part by mass of oxygen will combine with carbon $= \frac{27.3}{72.7} = 0.376$

Step-3.

To compare the masses of carbon which combine with the same mass of oxygen in both the oxides. The ratio of the masses of carbon that combine with the same mass of oxygen (1 part) is.

0.751 : 0.376 or 2 : 1

Since this is simple whole number ratio, so the above data illustrate the law of multiple proportions.

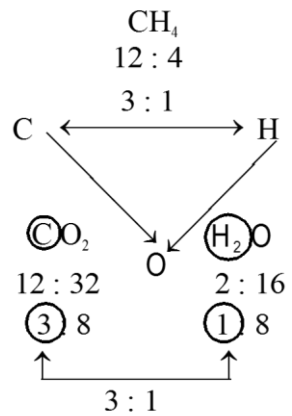
6. LAW OF RECIPROCAL PROPORTION (OR LAW OF EQUIVALENT WT.)

- It is given by Richter.

Statement

The ratio of the weights of two elements A and B which combine separately with a fixed weight of the third element C is either the same or simple ratio of the weights in which A and B combine directly with each other.

Ex.



- **Special Note:**

  This law is also called as law of equivalent wt. due to each element combined in their equivalent wt. ratio.

  $$E = \frac{M_W / \text{At.wt.}}{\text{V.F.}}$$

- **For Ions**

  V.F. = Total no. of positive charge

  or V.F. = Total no. of negative charge

Gay Lussac's Law of Gaseous Volumes The law developed by Gay Lussac in 1808 establishes that "the relationship between the volume of a gaseous reactant and a product can be represented by a simple whole number."

Avogadro Law

In 1811, Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

Avogadro's Hypothesis

According to Dalton's atomic theory, elements react with each other in the simple ratio of their atoms. Gay-Lussac proposed that gases combine in simple ratio of their volumes. In an attempt to correlate Dalton's atomic theory with Gay-Lussac law of gaseous volumes, Berzelius stated that under similar conditions of temperature and pressure, equal volume of all gases contain the same number of atoms. This hypothesis was subsequently found to be incorrect as it failed to interpret the experimental results and contradicted the very basic assumption of Dalton's atomic theory, i. e., an atom is indivisible. For example, the formation of hydrogen chloride from hydrogen and chlorine could not be explained on the basis of Berzelius hypothesis.

Hydrogen +	Chlorine =	Hydrogen chloride
I vol	1 vol	2 vol
nn atoms	n atoms	2π compound atoms
1 atom	1 atom	2 compound atoms
1/2 atom	1/2 atom	1 compound atom

for the formation of I compound atom of hydrogen chloride, ~ atom of hydrogen and ~ atom of chlorine is needed. In other words, each atom of hydrogen and chlorine has been divided which is against Dalton's atomic. theory. Thus, the hypothesis of Berzelius was discarded.

The Italian scientist, Amedeo Avogadro, in 1811, solved the above problem by proposing two types of particles from which whole of the matter is composed.

(i) Atom: The smallest particle of an element that can take part in chemical change but generally cannot exist freely as such.

(ii) Molecule: The smallest particle of a substance (element or compound) which has free or independent existence and possesses all characteristic properties of the substance. A molecule of an element is composed of like atoms while a molecule of a compound

contains fixed number of atoms of two or more different elements. A molecule may be broken down into its constituent atoms but the atom is indivisible during a chemical change.

Avogadro after making the above differentiation, presented a hypothesis known as Avogadro hypothesis which can be stated as follows:

"Under similar conditions of temperature and pressure, equal volumes of a D gases contain equal number of molecules."

Avogadro hypothesis explains successfully the formation of hydrogen chloride.

Hydrogen +	Chlorine =	Hydrogen chloride
I vol	1 vol	2 vol
n molecules	n molecules	2n molecules
1 molecule	1 molecule	2 molecules
$\frac{1}{2}$ molecule	$\frac{1}{2}$ molecule	1 molecule
1 atom	1 atom	1 molecule

(Both hydrogen and chlorine are diatomic in nature.)

Thus, the hypothesis explains that the molecules of reacting gases break up into constituent atoms during chemical change which then combine to form new molecules of the product or products.

Applications of Avogadro's hypothesis

(i) Atomicity: Atomicity means number of atoms present in one molecule of an elementary gas. Hydrogen, oxygen, nitrogen, chlorine, etc., are diatomic in nature. Noble gases are monoatomic while ozone is triatomic in nature. Avogadro's· hypothesis helps in determining the atomicity of elements.

(ii) Relationship between molecular mass and vapour density: The vapour density of any gas is the ratio of the densities of the· gas and hydrogen under similar conditions of temperature and pressure.

$$\text{Vapour Density (V.D.)} = \frac{\text{Density of gas}}{\text{Density of bydrogen}}$$

$$= \frac{\text{Mass of a certain volume of the gas}}{\text{Mass of same volume of bydrogen at}}$$

the same temp. and preserve

If n molecules are present in the given volume of a gas and hydrogen under similar conditions of temperature and pressure,

$$\text{V.D.} = \frac{\text{Mass of } n \text{ molecules of gas}}{\text{Mass of } n \text{ molecules of hydrogen}}$$

$$= \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 molecule of tydrogen}}$$

$$= \frac{\text{Molecular mass of gas}}{\text{Molecular mass of hydrogen}}$$

$$= \frac{\text{Mol. mass}}{2}$$

(since, mol. mass of hydrogen $= 2$)

Hence, 2x V.D.=Mol. Mass

This formula can be used for the detennination of molecular masses of volatile substances from vapour density.

Vapour density is measured mainly by two methods:

(a) Victor Meyer

(b) Duma's methods

(iii) Gram-molecular volume: I g mole of any gas occupies 22.4 litres or 22400 mL of volume at NTP or STP conditions.

The density of hydrogen at NTP is 0.00009 g mL^{-1}. Thus,

0.00009 g of hydrogen will occupy volume at NTP = 1 mL

$$\text{1 g of hydrogen occupies volume at NTP} = \frac{1}{0.00009} \text{ mL}$$

1g mole of hydrogen (2.016g) occupies volume at NTP

$= \frac{2.016}{0.00009} = 22400 \text{ mL} = 22.4 \text{ litre}$

According to Avogadro's hypothesis, equal volumes of different gases contain same number of molecules under similar conditions of temperature and pressure. Thus, 22.4 litter or 22400 mL of any gas at NTP will contain one gram mole or its molecular mass in grams.

Loschmidt number: Number of molecules in 1em3 or 1 mL of a gas at S.T.P. is known as Loschmidt number

$$\text{Loschmidt number} = \frac{6.023 \times 10^{23}}{22400}$$
$$= 2.68 \times 10^{18} \text{mL}^{-1}$$

(iv) Molecular formula: Avogadro's hypothesis helps in mding the molecular formulae of gases. Under similar conditions of temperature and pressure, 2 volumes of ozone after decomposition give 3 volumes of oxygen

$$\underset{\text{2 vol}}{\text{Ozone}} \xrightarrow{\text{Decimation}} \underset{\text{3 vol}}{\text{Oxygen}}$$

2 molecules	3 molecules
1 molecule	3/2 molecules
1 molecule	3 atoms

Thus, the formula of ozone is O_3,