LAWS OF CHEMICAL COMBINATION

A multitude of objects in our surroundings are visible and tangible. Matter is defined as anything that occupies space and has mass. John Dalton is credited with formulating the theory on the structure of matter, which is now recognized as Dalton's atomic theory.

To comprehend the composition of different compounds, it's essential to possess a theory that accommodates both qualitative and quantitative observations during chemical transformations.

- Antoine Lavoisier, John Dalton, and other scientists formulated specific laws related to the composition of matter and chemical reactions. These principles are collectively referred to as the laws of chemical combination.

1. The Law of Conservation of Mass

This principle, attributed to Lavoisier, is also known as the law of the indestructibility of matter. According to this law, the total mass of a system remains constant in all chemical changes; that is, mass is neither created nor destroyed during a chemical change.

In chemical changes, the conservation of total mass holds true. This means that the mass before the reaction is always equal to the mass after the reaction. All chemical reactions adhere to this law, making it the foundation for all quantitative work in chemistry.

Therefore, in a chemical change, the total masses of reactants are equal to the total masses of products. This relationship is valid when reactants are completely transformed into products. However, if the reacting materials are not entirely consumed, the relationship becomes:

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

In summary, the conservation of total mass is a fundamental aspect of chemical changes, ensuring that the mass before the reaction is always equal to the mass after the reaction.

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

(g) (g)

1 mole $\frac{1}{2}$ mole 1 mole

mass before the reaction = $1 \times 2 + 1/2 \times 32 = 18$ gm

mass after the reaction = $1 \times 18 = 18$ 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

2. **Law Of Constant or Definite Proportion**
Formulated by Proust in 1799, this law can be expressed as follows: A chemical compound consistently comprises the same elements combined in fixed proportions by mass. In other words, a chemical compound maintains a fixed composition that remains independent of its preparation method or the source from which it is derived. All chemical compounds exhibit a consistent composition, regardless of how they are prepared or their origin.

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,

$$\text{wt. of metal} = 1.80 \text{ g},$$
$$\text{wt. of oxygen} = (3.0 - 1.80) \text{ g} = 1.2 \text{ g}$$
$$\therefore \quad \frac{\text{wt.of metal}}{\text{wt.of oxygen}} = \frac{1.80 \text{ g}}{1.2 \text{ g}} = 1.5$$

In the second sample of the oxide,

$$\text{wt. of metal} = 1.50 \text{ g},$$
$$\text{wt. of oxygen} = (2.50 - 1.50) \text{ g} = 1 \text{ g}$$
$$\therefore \quad \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.50 \text{ g}}{1 \text{ g}} = 1.5$$

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.

3. **The Law of Multiple Proportion**
Proposed by Dalton in 1808, Dalton's Law states that if two elements combine to create more than one compound, the distinct masses of one element combining with a fixed mass of the other element exhibit a simple ratio to each other.
For instance, hydrogen and oxygen form two compounds: H_2O (water) and H_2O_2 (hydrogen peroxide). When an element combines with another element to produce multiple different compounds, the mass of one element, combining with a constant mass of the other, bears a simple ratio to each other.

Note: In this context, "simple ratio" refers to the ratio between small natural numbers, like 1:1, 1:2, 1:3. Subsequently, this straightforward ratio evolves into the element's valency and eventually its oxidation state.

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.

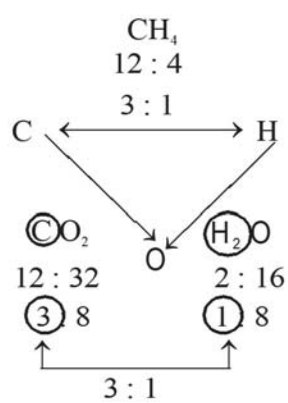
5. **Law Of Reciprocal Proportion (Or Law of Equivalent Wt.)**

It is given by Richter.

Statement

The ratio of the weights of two elements, A and B, combining individually with a constant weight of the third element, C, is either identical or a simple ratio to the weights in which A and B directly combine with each other.

Ex.



Special Note:

This principle is alternatively known as the law of equivalent weight because each element combines in its equivalent weight 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 formulated the Law of Gaseous Volumes in 1808, stating that "the correlation between the volume of a gaseous reactant and a product can be expressed by a straightforward whole number."

Avogadro Law

In 1811, Avogadro suggested that equal volumes of gases at identical temperature and pressure should encompass an equal number of molecules.

Avogadro's Hypothesis

Following Dalton's atomic theory, which posits that elements react in the simple ratio of their atoms, Gay-Lussac suggested that gases combine in the simple ratio of their volumes. In an effort to reconcile Dalton's atomic theory with Gay-Lussac's law of gaseous volumes, Berzelius asserted that, under equivalent conditions of temperature and pressure, all gases occupy equal volumes containing the same number of atoms. However, this hypothesis proved incorrect as it did not align with experimental findings and contradicted the fundamental assumption of Dalton's atomic theory—that 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 creation of one molecule of hydrogen chloride, approximately one atom of hydrogen and one atom of chlorine are required. This implies that each atom of hydrogen and chlorine has undergone division, contradicting Dalton's atomic theory. Consequently, Berzelius's hypothesis was rejected.

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 tiniest constituent of an element capable of participating in chemical reactions, typically unable to exist independently.

(ii) **Molecule:** The smallest unit of a substance (element or compound) with independent existence and displaying all the characteristic properties of the substance. An element's molecule consists of identical atoms, while a compound's molecule contains a fixed number of atoms from two or more different elements. Although a molecule can be broken down into its constituent atoms, the atom remains indivisible during a chemical change. Following this distinction, Avogadro proposed a hypothesis, often referred to as Avogadro's hypothesis, which states: "Under similar conditions of temperature and pressure, equal volumes of gases contain an 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.)

Hence, the hypothesis clarifies that, during a chemical change, the molecules of the reacting gases disintegrate into individual atoms, which subsequently combine to create new molecules of the product or products.

Applications of Avogadro's hypothesis

(i) Atomicity

Atomicity refers to the number of atoms found in a single molecule of an elementary gas. Gases such as hydrogen, oxygen, nitrogen, chlorine, etc., exhibit diatomic nature. Noble gases exist as monoatomic, while ozone possesses a triatomic nature. Avogadro's hypothesis plays a crucial role in ascertaining the atomicity of elements.

(ii) Relationship between molecular mass and vapour density

The vapor density of a gas is the ratio of the gas density to the density of hydrogen, both measured 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,

$$2\text{x V.D.} = \text{Mol. Mass}$$

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 liters 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

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

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

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

In accordance with Avogadro's hypothesis, equal volumes of diverse gases encompass an identical number of molecules when subjected to comparable conditions of temperature and pressure. Consequently, 22.4 liters or 22,400 mL of any gas at NTP (Normal Temperature and Pressure) will contain one gram mole or the molecular mass expressed in grams.

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

Loschmidt number: The quantity of molecules presents in 1 cm^3 or 1 mL of a gas at Standard Temperature and Pressure (S.T.P.) is referred to as the Loschmidt number.

(iv) Molecular formula

Avogadro's hypothesis aids in determining the molecular formulas of gases. In analogous conditions of temperature and pressure, the decomposition of 2 volumes of ozone results in the formation of 3 volumes of oxygen.

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

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

Thus, the formula of ozone is O_3,

6. Gay Lussac's Law of Combining Volumes

In accordance with this scientific principle, when gases undergo chemical reactions, their volumes exhibit a consistent pattern: they consistently maintain simple ratios to one another and to the volume of the resulting products, particularly when these products are also in the gaseous state. It's crucial to note that this observation holds true under the condition that all volume measurements are conducted in a uniform environment characterized by similar conditions of temperature and pressure. This principle provides valuable insights into the relationship between the volumes of reacting gases and the volumes of the resulting products, offering a foundational understanding for the study of gas reactions under controlled experimental conditions.

The balanced chemical equation for the reaction $H_2(g) + Cl_2(g) \rightarrow 2HCl$ indicates that one volume of hydrogen gas (H_2) reacts with one volume of chlorine gas ($2Cl_2$) to produce two volumes of hydrogen chloride gas (2HCl). This relationship underscores the principle described by Avogadro's Law, where gases under similar conditions of temperature and pressure react in volumes that bear simple ratios to one another. In this specific reaction, the simple ratio is 1:1:2 for H_2, Cl_2, and HCl respectively.