LECT-1

Antoine Laurent Lavoisier is called the father of chemistry. Chemistry is the branch of science which deals with the composition, structure, and properties of matter. These aspects can be best described and understood in terms of basic constituents of matter: atoms and molecules. That is why chemistry is called the science of atoms and molecules.

Importance of Chemistry:

Chemistry also plays an important role in daily life. Chemical principles are important in diverse areas, such as weather patterns, functioning of the brain, and operation of a computer. Chemical industries manufacturing fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys, and other inorganic and organic chemicals.

Chemistry plays an important role to fulfill human needs for food, health care products, a variety of fertilizers, varieties of pesticides, insecticides, and other materials aimed at improving the quality of life. Similarly, many life-saving drugs such as cisplatin and taxol, are effective in cancer therapy and AZT (Azidothymidine) is used for helping AIDS victims.

Branches of Chemistry

- **Organic Chemistry** -This branch deals with the study of carbon compounds especially hydrocarbons and their derivatives.

- **Inorganic Chemistry-** This branch deals with the study of compounds of all other elements except carbon. It largely concerns itself with the study of minerals found in the earth's crust.

Physical Chemistry-The explanation of fundamental principles governing various chemical phenomena is the main concern of this branch. It is concerned with laws and theories of the different branches of chemistry.

- **Industrial Chemistry-**

The chemistry involved in industrial processes is studied under this branch.

- **Analytical Chemistry-**

This branch deals with the qualitative and quantitative analysis of various substances.

Biochemistry- This branch deals with the chemical changes going on in the bodies of living organisms; plants and animals.

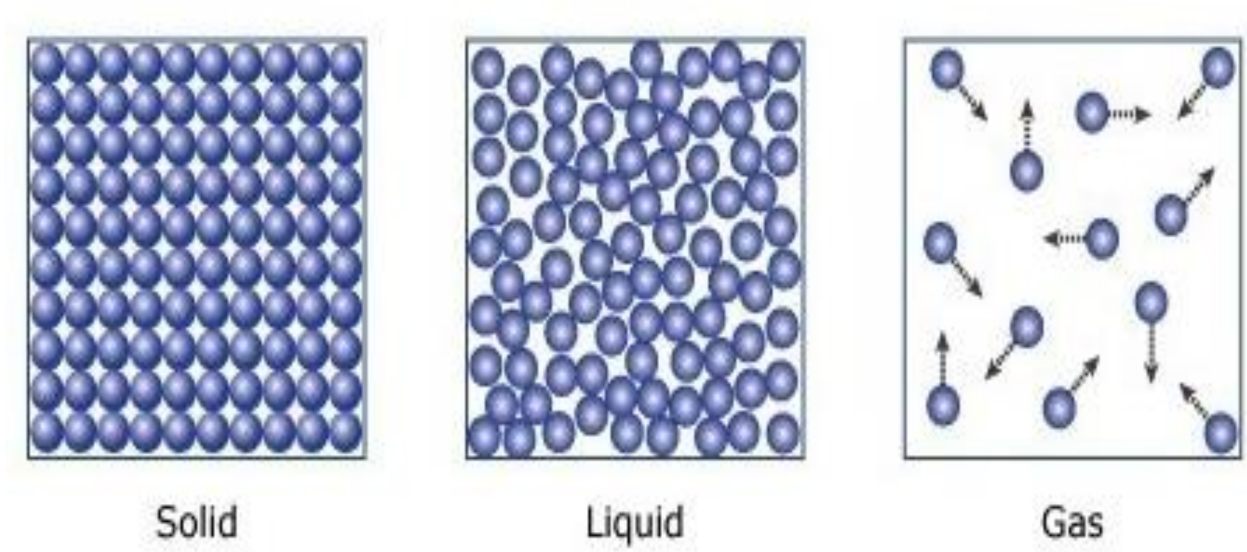
- **Nuclear Chemistry-** Nuclear reactions, such as nuclear fission, nuclear fusion, transmutation processes, etc. are studied under this branch.

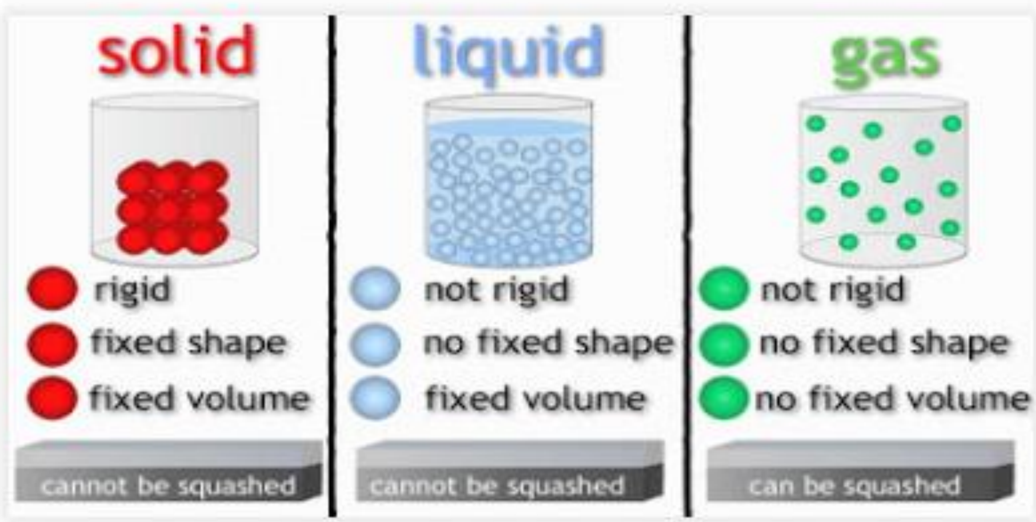
Nature of Matter

Anything which has mass and occupies space is called matter.

Everything around us, for example, book, pen, pencil, water, air, all living beings, etc. are composed of matter. They have mass and they occupy space.

Matter can exist in three physical states *viz. solid, liquid, and gas.*
The constituent particles of matter in these three states can be represented as:





In solids, these particles are held very close to each other in an orderly fashion and there is not much freedom of movement.
In liquids, the particles are close to each other but they can move around.
However, **in gases**, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.
Because of such arrangement of particles, different states of matter exhibit the following characteristics:
(i) Solids have *definite volume and definite shape.*

(ii) Liquids have *definite volume but not a definite shape. They take the shape of the* container in which they are placed.

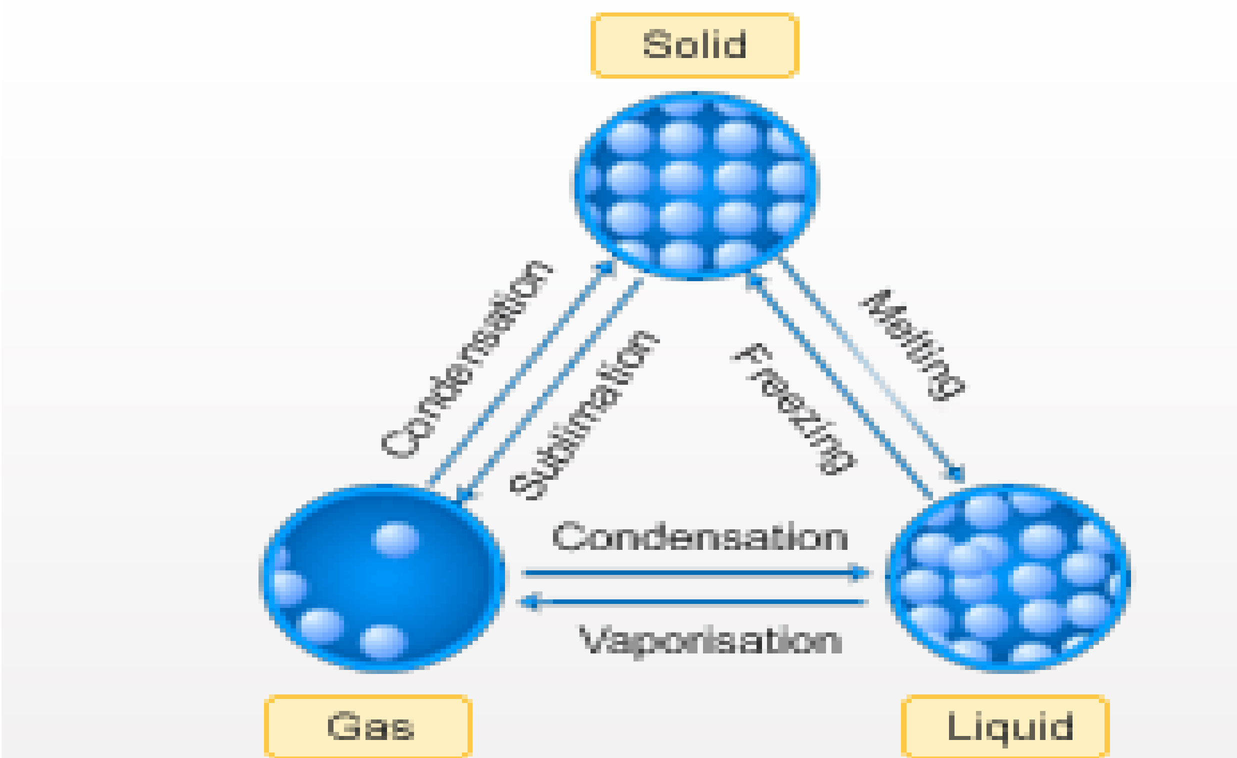
(iii) Gases have *neither definite volume nor definite shape. They completely occupy the* container in which they are placed.

These three states of matter are interconvertible by changing the conditions of temperature and pressure

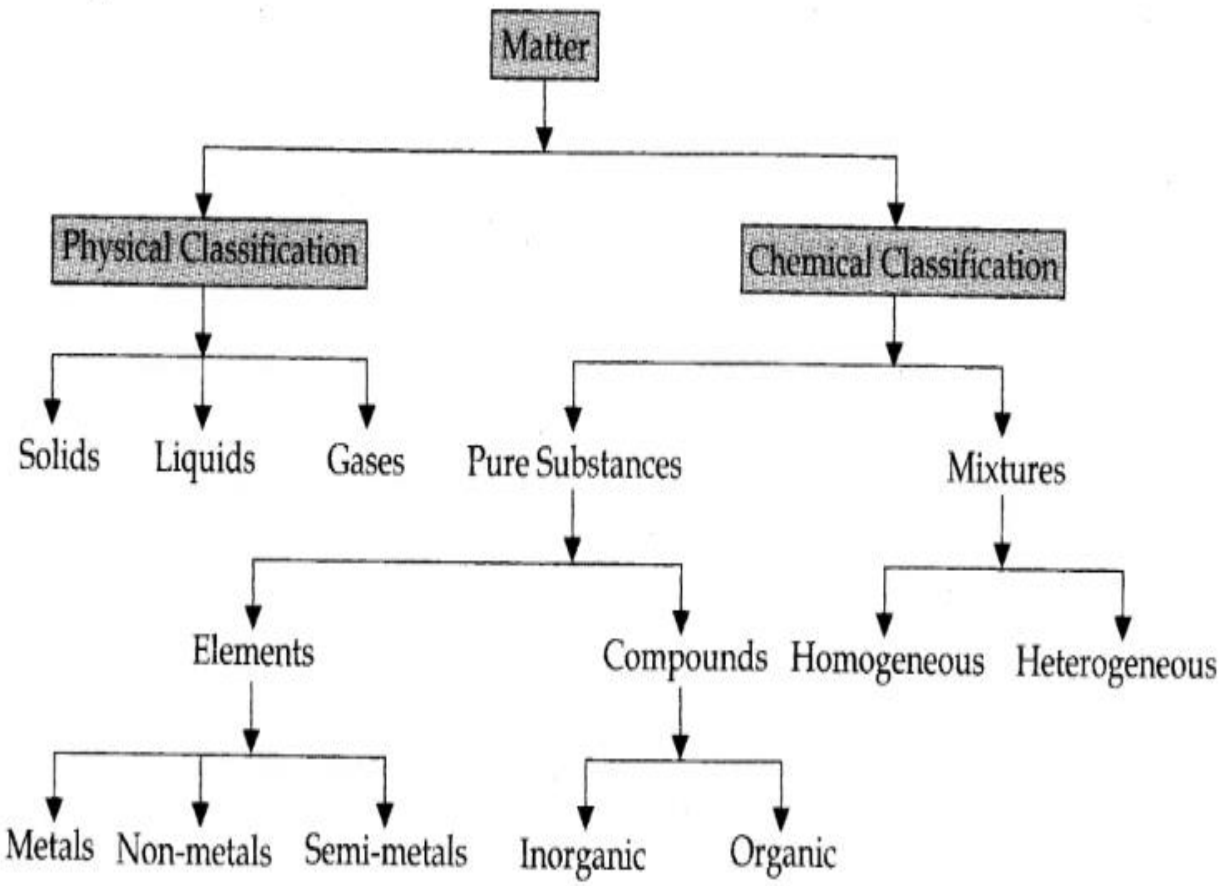
$$\text{Solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Gas}$$

On heating, a solid usually changes to a liquid, and the liquid on further heating changes to the gaseous (or vapor) state.

In the reverse process, a gas on cooling liquifies to the liquid, and the liquid on further cooling freezes to the solid.



At the macroscopic or bulk level, matter can be classified as mixtures or pure substances. These can be further sub-divided as:



Many of the substances present around us are mixtures. For example, sugar solutions in water, air, tea, etc., are all mixtures. A mixture contains two or more substances present in it (in any ratio) which are called its components.

A mixture may be homogeneous or heterogeneous.

In a **homogeneous mixture,** the components completely mix and its composition is uniform throughout.

Ex: Sugar solution and air are thus, examples of homogeneous mixtures.

In **heterogeneous mixtures**, the composition is not uniform throughout.

For example, the mixtures of salt and sugar, grains and pulses along with some dirt (often stone) pieces, are heterogeneous.

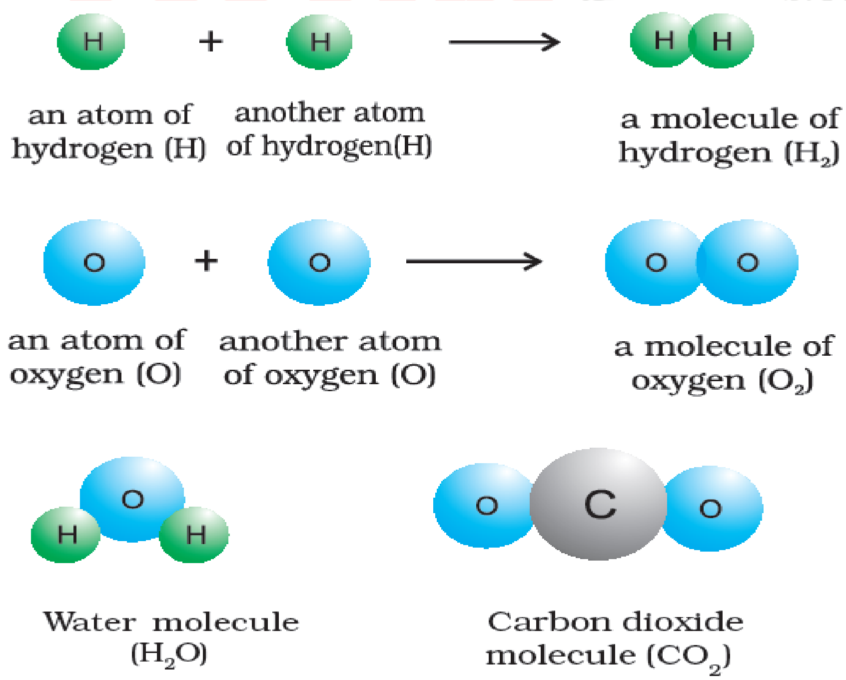
Pure substances: These are different from the mixtures. They have fixed composition, whereas mixtures may contain the components in any ratio and their composition is variable.

Copper, silver, gold, water, glucose are some examples of pure substances. Glucose contains carbon, hydrogen, and oxygen in a fixed ratio and thus like all other pure substances have a fixed composition. Also, the constituents of pure substances cannot be separated by simple physical methods.

Pure substances can be further classified into elements and compounds. An element consists of only one type of particle. These particles may be atoms or molecules.
Sodium, copper, silver, hydrogen, oxygen, etc. are some examples of elements. They all contain atoms of one type.
However, the atoms of different elements are different. Some elements such as sodium or copper, contain single atoms held together as their constituent particles whereas, in some others, two or more atoms combine to give molecules of the element. Thus, hydrogen, nitrogen, and oxygen gases consist of molecules in which two atoms combine to give their respective molecules.
When two or more atoms of different elements combine, the molecule of a compound is obtained.
Examples of some compounds are water, ammonia, carbon dioxide, sugar, etc.

Atoms of different elements

Representation of atoms and molecules



The atoms of different elements are present in a compound in a fixed and definite ratio and this ratio is characteristic of a particular compound. Also, the properties of a compound are different from those of its constituent elements. For example, hydrogen and oxygen are gases whereas the compound is formed by their combination *i.e., water is a liquid.*

• **Differences between Compounds and Mixtures**

Compounds

1. In a compound, two or more elements are combined chemically.
2. In a compound, the elements are present in the fixed ratio by mass. This ratio cannot change.
3. CompoUnds are always homogeneous i.e., they havethe same composition throughout.
4 In a compound, constituents cannot be separated by physical methods
5. In a compound, the constituents lose their identities i.e., i compound does not show the characteristics of the constituting elements.

Mixtures

1. In a mixture, or more elements or compounds are simply mixed and not combined chemically.
2. In a mixture the constituents are not present in fixed ratio. It can vary
3. Mixtures may be either homogeneous or heterogeneous in nature.
4. Constituents of mixtures can be separated by physical methods.
5, In a mixture, the constituents do not lose their identities i.e., a mixture shows the characteristics of all the constituents .

1. In what way iodine can be separated from a mixture of potassium chloride and iodine?

2. A mixture of salt and sugar is known as

a) Homogeneous mixture b) Heterogeneous mixture
c) Mixture of Components d) None of the above

3. **Name the particles which make up matter?**

a) Non metals b) Metals
c) Metalloids d) Atoms

4. Which of the following is the property of Solid?

a) Can be compressed b) Have a definite shape
c) Have low-density d) Intermolecular force is less

5. Which of the following is a property of diffusion?

a) Slowest in liquids
b) Fastest in gases
c) Based on the motion of particles
d) All of the above.

6. Which one of the following substances is a compound?

a) Iron b) Hydrogen c) Oxygen d)Sodium fluoride

7. Which of the following is Molecule?

a) Na b) Fe c) NaCl d) Ne

8. Which of the following is an element?

a) Oxygen b) Ozone c) Nitrogen d) all of these

9. Which of the following is a heteroatomic molecule?
   a) Ozone b) Sulphur c) Water d) Dinitrgen
10. Which is not homogeneous?
   a) Air b) Storm c) Tap water d) Both (b) & (c)
11. What are basic radicals, acid radicals, and free radicals?
12. Write the formula of the following compounds
   a) Sod.sulphate b) Sod.Phosphate c) Aluminium carbide
   d) Aluminium sulphate e) Calcium carbide f) Sod.acetate
   g) Cal.acetate h) Sod.oxalate i) Cal.phosphate j) Nitrous acid
   k) Perchloric acid l) Iodic acid m) Sulphurous acid
   n) Platinum chloride

LECT-2

LAWS OF CHEMICAL COMBINATIONS

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

1. Law of Conservation of Mass

It states that matter can neither be created nor destroyed.
This law was put forth by Antoine Lavoisier in 1789.
In a chemical reaction, the total mass of reactants equals the total mass of products.

2. Law of Definite Proportions

This law was given by, a French chemist, Joseph Proust. He stated that a given compound always contains the same proportion of elements by weight.
In water, H:O is always 1:8. Whatever may be the source of water, the ratio between hydrogen & oxygen is always fixed.

Law of Multiple Proportions

This law was proposed by Dalton in 1803. According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.
For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen + Oxygen → Water
2g 16g 18g

Hydrogen + Oxygen → Hydrogen Peroxide
2g 32g 34g

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2

Explanation: Let two different elements C & O combine to form CO and CO_2. The different masses of oxygen i.e 16 and 32 combined with a fixed mass of Carbon(12). The ratio of different masses of oxygen bears a simple whole-number ratio to one another.

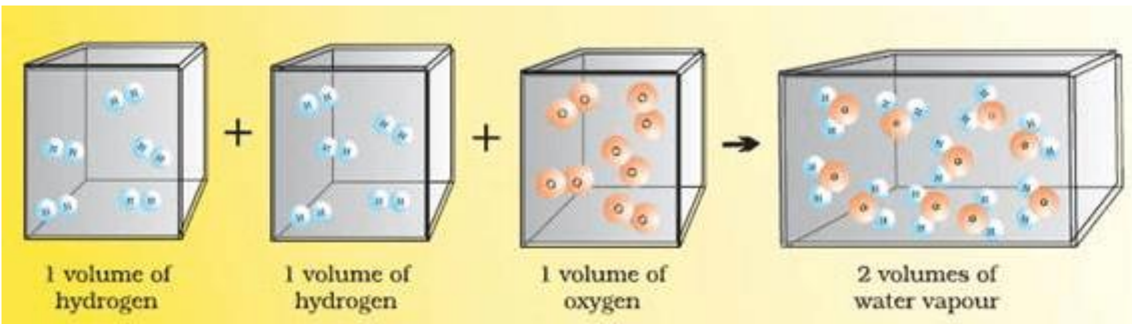
In CO, C:O = 12:16 = 3:4 -------------(i)

In CO_2, C:O = 12:32= 3: 8 ------------(ii)

Here mass of Carbon is kept constant and the ratio of different masses of oxygen is 16:32=1:2 which is a simple whole-number ratio.

Gay Lussac's Law of Gaseous Volumes

This law was given by Gay Lussac in 1808. He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at the same temperature and pressure.



Thus, 100 mL of hydrogen combines with 50 mL of oxygen to give 100 mL of water vapor.

Hydrogen + Oxygen → Water

100 ml 50 ml 100 ml

Thus, the volumes of hydrogen and oxygen which combine (i.e. 100 ml and 50 ml) bear a simple ratio of 2:1.

H_2 (g) + Cl_2(g) → 2HCl (g)

1 vol. 1 vol. 2 vol.

The ratio of volumes is 1:1:2 which is a simple whole-number ratio.

1. A water sample from a lake, ocean, rain, or pond must have _____ proportions of hydrogen to oxygen.

2. A chemical equation is balanced according to the law of ----------

3. 5.2 g of $CaCO_3$ when heated produced 1.99 g of Carbon dioxide and the residue (CaO) left behind weighs 3.2g. Show that these results illustrate the law of conservation of mass.

4. The law of multiple proportion is illustrated by the pair of compounds:

a. sodium chloride and sodium bromide

b. water and heavy water

c. sulfur dioxide and sulfur trioxide

d. magnesium hydroxide and magnesium oxide

5. 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.

LECT-3

Topic: Dalton's atomic theory and Avogadro's law Atomic and molecular masses, average atomic mass, and formula mass:

The matter is composed of small indivisible particles called *'a-tomio' (meaning -indivisible),*
In 1808, Dalton proposed the following :

1. All matter consists of indivisible particles called atoms.
2. Atoms of the same element are similar in shape and mass but differ from the atoms of other elements.
3. Atoms cannot be created or destroyed.
4. Atoms of different elements may combine in fixed, simple, whole-number ratios to form compound atoms.
5. Atoms of the same element can combine in more than one ratio to form two or more compounds.
6. The atom is the smallest unit of matter that can take part in a chemical reaction.

Drawbacks of Dalton's Atomic Theory

1. The indivisibility of an atom was proved wrong: An atom can be further subdivided into protons, neutrons, and electrons. However, an atom is the smallest particle that takes part in chemical reactions.
2. According to Dalton, the atoms of the same element are similar in all respects. However, atoms of some elements vary in their masses and densities. These atoms of different masses are called isotopes. For example, chlorine has two isotopes with mass numbers 35 and 37.
3. Dalton also claimed that atoms of different elements are different in all respects. This has been proven wrong in certain cases: argon and calcium atoms each have an atomic mass of 40 amu. These atoms are known as isobars.
4. According to Dalton, atoms of different elements combine in simple whole-number ratios to form compounds. This is not observed in complex organic compounds like sugar ($C_{12}H_{22}O_{11}$).
5. The theory fails to explain the existence of allotropes; it does not account for differences in properties of charcoal, graphite, diamond.

The Gay-Lussac's law was explained properly by the work of Avogadro

Avogadro 's Law: Avogadro proposed that ''equal volumes of gases at the same temperature and pressure should contain an equal number of molecules.

Avogadro made a distinction between atoms and molecules which is quite understandable in the present times. If we consider again the reaction of hydrogen and oxygen to produce water, we see that two volumes of hydrogen combine with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen.

Application of Avogadro's law

- To determine the atomicity of elementary gases.

- To prove Mol.mass= 2 x V.D
- To prove molar volume is 22.4 litres.
  - **To prove atomicity**

Hydrogen + Oxygen → Water vapour
2vol. 1vol. 2vol.
Aplying Avogadro's law
2n molecules n molecules 2n molecules
2 molecules 1 molecules 2 molecules
1 molecule ½ molecule 1 molecule
½ molecule of oxygen = 1 atom of oxygen
1 molecule of oxygen = 2 atoms of oxygen
Thus atomicity of oxygen= 2

ATOMIC MASS or ATOMIC WEIGHT

Introduction: As atoms are too small to be isolated, therefore it is not possible to determine the actual mass of an atom by direct weighing. Because of this problem, some indirect methods were suggested to find out the mass of atoms. One of the possible way which is suggested as,

A sample of an element is taken and weighed and then by dividing the total mass of the sample by the total number of atoms in the sample, the mass of an atom of the element could be found out. However, there was a difficulty to calculate the number of atoms in a sample by counting. This problem was solved by Avogadro's law which states that " Equal volume of different gases under STP contains the equal number of molecules. If an equal volume of two different gases are taken under STP and then weighed, the ratio of their masses will be equal to the ratio of masses of their single molecule.

Ex: Ratio of masses of an equal volume of H_2 and O_2 found as 1:16. It means that a molecule of oxygen is 16 times heavier than a molecule of H_2. So It is believed that an atom of oxygen is 16 times heavier than an atom of hydrogen.

In the beginning, the atomic masses of all elements were obtained by comparing with the mass of hydrogen taken as 1. Hence the relative mass of an atom of oxygen is 16 units concerning hydrogen atom as standard scale(because it is the lightest element). But by doing so, the atomic mass of most of the elements came out to be fractional. Hence the reference was changed to oxygen taken as 16 units. But the relative atomic mass of other elements was found to more close to whole numbers.

In 1961, the international union of chemists selected the most stable isotope of carbon-12 as the standard scale for the comparison of the atomic mass of various elements. The atomic mass of an element is the number of times an atom of that element is heavier than an atom of carbon taken as 12.

The scale in which the relative atomic mass of different atoms is expressed is called the atomic mass unit scale or a.m.u scale. One a.m.u is equal to 1/12th of the mass of an atom of the carbon-12 isotope.

$$\text{Atomic mass} = \frac{\text{Mass of an atom of the element}}{\text{1/12 th of mass of an atom of C-12}}$$

It has been found that 1 a.m.u = $1.66056\ \times 10^{-24}$ g

Mass of an atom of hydrogen = $1.6736\times\ 10^{-24\mathrm{g}}$

$$1\,\text{a.m.u} = 1/12\text{th of mass of one atom of } C^{12} \text{ isotope}$$

$$= \frac{1}{12}\times\frac{12}{6.023\times10^{23}} = \frac{1}{6.023\times10^{23}} = 1.66\times10^{\ 24}\,\text{gm}$$

Gram Atomic Mass: Gram atomic mass or molar mass of an element is the mass of 1 mole of atoms of that element or it is the relative atomic mass expressed in gram.

The gram atomic mass of O = Mass of 1 mole of O = Molar mass of O = Mass of gram atom of O = 16 gram.

Kg Atomic mass or Killo Molar mass: Mass of the one-kilo mole of atoms of an element is known as Kg atomic mass or kilo molar mass of the element. It is also called relative atomic mass in Kg.

Ex: Kg atomic mass of O = Mass of 1000 mole of O = Mass of Kg atom of O = Killo molar mass of O = 16 kg.

Average atomic mass: In the case of isotopes, the atomic mass of the element is taken as the average value.

Ex: Ordinary Chlorine is found as a mixture of two isotopes of atomic masses 35 u and 37 u in a 3: 1 ratio.

$$\text{Average atomic mass} = \frac{35\times3+37\times1}{3+1} = 35.5\,\text{u}$$

Hence the atomic mass of an element is accurately defined as the atomic mass of an element is the average relative mass of an atom compared with an atom of C-12.

From the percentage of the abundance of different isotopes, the average atomic mass is calculated as:

$$\text{Avg. atomic mass} = \frac{\Sigma P_i A_i}{100} \quad \text{Where } P_i = \%\text{ of abundance of isotopes}$$

$$A_i = \text{Isotopic Mass}$$

Ex: Relative abundances of C^{12}, C^{13}& C^{14} are 98.892 %, 1.108 % and 2 x 10^{-10} % with isotopic mass 12, 13.0035 and 14.00317 respectvely.

$$\text{Avg. atomic mass of C} = \frac{98.892\times 12 + 1.108\times 13.00335 + 2\times 10^{-10}\times 14.00317}{100}$$
$$= 12.011\text{u}$$

Molecular Mass

Molecular mass is the sum of the atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together. For example, the molecular mass of methane which contains one carbon atom and four hydrogen atoms can be obtained as follows :

The molecular mass of methane,

(CH_4) = (12.011 u) + 4 (1.008 u)

= 16.043 u

Similarly, the molecular mass of water (H_2O)

= 2 × atomic mass of hydrogen + 1 × atomic

mass of oxygen

= 2 (1.008 u) + 16.00 u

= 18.02u

1. Calculate Molecular mass of Glucose ($C_6H_{12}O_6$)

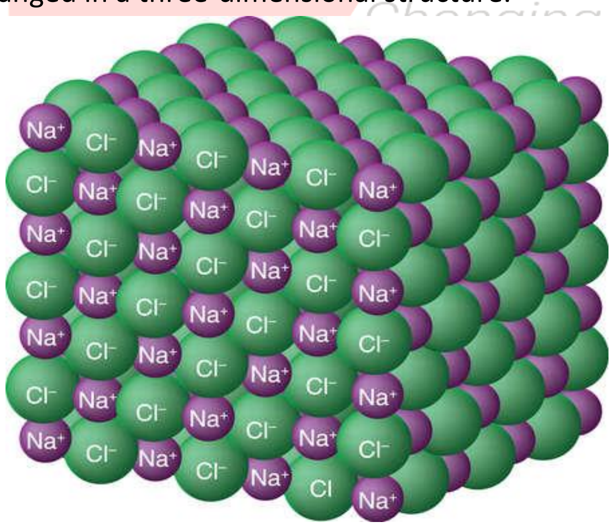
Solution: Molecular mass of glucose ($C_6H_{12}O_6$)

= 6(12.011 u) + 12(1.008 u) + 6(16.00 u)

= (72.066 u) + (12.096 u) + (96.00 u)

= 180.162 u

Formula Mass: Some substances such as sodium chloride do not contain discrete molecules as their constituent units. In such compounds, positive (sodium) and negative (chloride) entities are arranged in a three-dimensional structure.



It may be noted that in sodium chloride, one Na^+ is surrounded by six Cl^- and *vice-versa.* A formula such as NaCl is used to calculate the formula mass instead of molecular mass as in the solid-state sodium chloride does not exist as a single entity. Thus, formula mass of sodium chloride = atomic mass of sodium + atomic mass of Chlorine = 23.0 u + 35.5 u = 58.5 u.

LECT-5

MOLE CONCEPT

Atoms and molecules are extremely small in size and their numbers in even a small amount of any substance are very large. To handle such large numbers, a unit of similar magnitude is required. Just as we denote one dozen for 12 items, the score for 20 items, gross for 144 items, we use the idea of the mole to count entities at the microscopic level (i.e. atoms/molecules/ particles, electrons, ions, etc). In the SI system, mole (symbol, mol) was introduced as the seventh base quantity for the amount of a substance.

One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the 12C isotope.

It may be emphasized that the mole of a substance always contains the same number of entities, no matter what the substance may be. To determine this number precisely, the mass of a carbon–12 atom was determined by a mass spectrometer and found to be equal to 1.992648×10^{-23} g. Knowing that one mole of carbon weighs 12 g, the number of atoms in it is equal to

$$\frac{12g/mol\ {}^{12}c}{1.992648\times10^{-23}\,g/{}^{12}C-atom}$$

$$=6.0221367\times10^{23}\ atoms/mol$$

This number of entities in 1 mol is so important that it is given a separate name and symbol. It is known as the 'Avogadro constant', denoted by N_A in honor of Amedeo Avogadro. To appreciate the largeness of this number, let us write it with all the zeroes without using any powers of ten.

602213670000000000000000

Hence so many entities (atoms, molecules, or any other particle) constitute one mole of a particular substance.

therefore, say that 1 mol of hydrogen atoms = 6.022×10^{23} atoms

1 mol of water molecules = 6.022×10^{23} water molecules

1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride

The mass of one mole of a substance in grams is called its gram molar mass. The molar mass in grams is numerically equal to atomic/molecular/ formula mass in u.

The molar mass of water = 18.02 g mol^{-1}. The molar mass of sodium chloride = 58.5 g mol^{-1}

Important formulae of mole concept:

Mole and Gram Atomic Mass: One mole of atoms = 6.022×10^{23} atoms = Gram atomic mass of an element = 1 gram atom of the element
Mole and Gram Molecular Mass: One mole of molecules = 6.022×10^{23} molecules = Gram molecular mass = 1 gram molecule of the compound
Mole in terms of volume: One mole of a gas = 22.4 litres at S.T.P.

Moles of an element = Mass of the element Atomic mass or GAW	Moles of a compound = Mass of the compound Molecular mass or GMW
Mass of one atom = Atomic Mass or GAW 6.022×10^{23}	Mass of one molecule = Molecular Mass or GAW 6.022×10^{23}
Number of molecules = Moles × 6.022 $\times 10^{23}$	Number of atoms = Moles × 6.022×10^{23}

$$\text{No.of moles} = \frac{\text{Mass of the element}}{\text{Gram atomic mass}} \quad \text{(For elements)}$$

$$\text{No.of moles} = \frac{\text{Mass of the compound}}{\text{Gram molecular mass}} \quad \text{(For compounds)}$$

$$\text{No.of moles} = \frac{\text{Volume of the gas}}{\text{Molar volume (22.4 L)}} \quad \text{(For gases)}$$

$$\text{Mass of an atom} = \frac{\text{Gram atomic mass}}{6.022 \times 10^{23}}$$

$$\text{Mass of a molecule} = \frac{\text{Gram molecular mass}}{6.022 \times 10^{23}}$$

$$\text{No.of atoms in given sample of an element} = \text{No.of moles} \times 6.022 \times 10^{23}$$

$$\text{No.of atoms in given sample of a compound} = \text{Atomicity} \times \text{No.of moles of the compound} \times 6.022 \times 10^{23}$$

LECT-6

Numerical related to Mole Concept:

1. Calculate the mass of 6.022×10^{23} molecule of Calcium carbonate ($CaCO_3$).
2. How many molecules of water are there in 54 g of H_2O?
3. Calculate the mass of 6.022 molecules of NH_4Cl?
4. How many atoms of hydrogen are there in 36 g of NH_3?
5. Calculate the number of molecules in 22.4 liters of CH_4 gas at NTP.
6. The atomic masses of two elements (P and Q) are 20 and 40 respectively. x g of P contains y atoms, how many atoms are present in 2x g of Q?
7. Calculate the mass of 12.044×10^{23} carbon atoms.
8. Calculate the number of oxygen atoms in 1 mole of O_2.
9. Calculate the number of Cu atoms in 0.635g of Cu.
10. Calculate the number of molecules in 11.2 liters of SO_2 gas at NTP.
11. An atom of some element X weighs 6.644×10^{-23} g. Calculate the number of gram-atoms in 40 kg of it.
12. Calculate the volume occupied by the 1-mole atom of (i) monoatomic gas, and (ii) diatomic gas at NTP.
13. Calculate the volume of 20g H_2 at NTP.
14. What is the volume occupied by 6.022×10^{23}molecules of any gas at NTP?
15. Calculate the number of atoms in 5.6 liters of an (i) monoatomic, and (ii) diatomic gas at NTP.
16. Calculate the number of atoms in 100 u of He.
17. If a mole were to contain 1×10^{24} particles, what would be the mass of (i) one mole of oxygen, and (ii) a single oxygen molecule?
18. Calculate the standard molar volume of oxygen gas. The density of O_2 gas at NTP is 1.429g/L.
19. Calculate the mass of 1 mole He gas. The density of He gas at NTP is 0.1784g/L.
20. A metal M of atomic mass 54.94 has a density of 7.42g/cc. Calculate the apparent volume occupied by one atom of the metal.
21. Calculate the number of moles, and the number of atoms of H, S, and O in 5 moles of H_2SO_4.
22. Calculate the number of oxygen atoms and their mass in 50 g of $CaCO_3$.
23. Calculate the number of atoms of each element in 122.5 g of $KClO_3$.
24. Calculate the total number of electrons present in 1.6 g of CH_4.
25. Find the charge of 1 g-ion of N^{3-} in *Coulombs.*
26. Equal masses of oxygen, hydrogen, and methane are taken in a container in identical conditions. Find the ratio of the volumes of the gases.

LECT-7

PERCENTAGE COMPOSITION

Let us understand it by taking the example of water (H_2O). Since water contains hydrogen and oxygen, the percentage composition of both these elements can be calculated as follows :

Mass % of an element =

$$\text{Mass \% of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$$

$$\text{Molar mass of water} = 18.02 \text{ g}$$

$$\text{Mass \% of hydrogen} = \frac{2 \times 1.008}{18.02} \times 100 = 11.18$$

$$\text{Mass \% of oxygen} = \frac{16.00}{18.02} \times 100 = 88.79$$

Let us take one more example. What is the percentage of carbon, hydrogen and oxygen in ethanol?

Molecular formula of ethanol is : C_2H_5OH

Molar mass of ethanol is : $(2 \times 12.01 + 6 \times 1.008 + 16.00)$ g $= 46.068$ g

Mass per cent of carbon

$$= \frac{24.02 \text{ g}}{46.068 \text{ g}} \times 100 = 52.14\%$$

Mass per cent of hydrogen

$$= \frac{6.048 \text{ g}}{46.068 \text{ g}} \times 100 = 13.13\%$$

Mass per cent of oxygen

$$= \frac{16.00 \text{ g}}{46.068 \text{ g}} \times 100 = 34.73\%$$

Empirical Formula and Molecular Formula

Molecular formula: It shows the exact number of different types of atoms present in a molecule of a compound. If the mass percent of various elements present in a compound is known, its empirical formula can be determined. The molecular formula can further be obtained if the molar mass is known.

Ex: M.F of Glucose is $C_6H_{12}O_6$. It represents the actual number of atoms present in the molecule.

Empirical Formula: An empirical formula represents the simplest whole-number ratio of various atoms present in a compound.

Ex: Empirical formula of Glucose($C_6H_{12}O_6$) is CH_2O (C:H: O=1:2:1).

The relation between Molecular formula and Empirical formula:

E.F x n = M.F .

Thus, Empirical formula mass x n = Molecular mass

Ex: E.F (CH_2O) x 6 = $C_6H_{12}O_6$ is M.F of Glucose.

E.F mass = 30 and M.M= 180

Thus, 30 x 6 = 180

Problem:

A compound contains 4.07 % hydrogen, 24.27 % carbon, and 71.65 % chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formula?

Solution

Step 1.Conversion of a mass percent to grams.

Since we are having a mass percent, it is convenient to use 100 g of the compound as the starting material. Thus, in the 100 g sample of the above compound, 4.07g hydrogen is present, 24.27g
carbon is present and 71.65 g chlorine is present.

Step 2. Convert into number moles of each element

Divide the masses obtained above by the respective atomic masses of various elements.

$$\text{Moles of hydrogen} = \frac{4.07\,\text{g}}{1.008\,\text{g}} = 4.04$$

$$\text{Moles of carbon} = \frac{24.27\,\text{g}}{12.01\text{g}} = 2.021$$

$$\text{Moles of chlorine} = \frac{71.65\text{g}}{35.453\,\text{g}} = 2.021$$

Step 3. Divide the mole value obtained above by the smallest number: Since 2.021 is the smallest value, division by it gives a ratio of 2:1:1 for H:Cl: In case the ratios are not whole numbers, then they may be converted into whole numbers by multiplying by the suitable coefficient.

Step 4. Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

CH_2Cl is, thus, the empirical formula of the above compound.

Step 5. Writing molecular formula

(a) Determine empirical formula mass Add the atomic masses of various atoms present in the empirical formula.

For CH_2Cl, empirical formula mass is 12.01 + 2 × 1.008 + 35.453 = 49.48 g

(b) Divide Molar mass by empirical formula mass

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96\,g}{49.48\,g}$$

$$= 2 = (n)$$

(c) Multiply empirical formula by n obtained above to get the molecular formula

Empirical formula = CH_2Cl, $n = 2$. Hence molecular formula is $C_2H_4Cl_2$.

LECT-8

1. Determine the empirical formula for compounds with the following percent compositions:

(a)15.8% carbon and 84.2% sulfur
(b) 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen

2. Determine the empirical and molecular formula for chrysotile asbestos. Chrysotile has the following percent composition: 28.03% Mg, 21.60% Si, 1.16% H, and 49.21% O. The molar mass for chrysotile is 520.8 g/mol.

3. A major textile dye manufacturer developed a new yellow dye. The dye has a percent composition of 75.95% C, 17.72% N, and 6.33% H by mass with a molar mass of about 240 g/mol. Determine the molecular formula of the dye.

4. Polymers are large molecules composed of simple units repeated many times. Thus, they often have relatively simple empirical formula. Calculate the empirical formula of the following polymers:

(a) Lucite (Plexiglass); 59.9% C, 8.06% H, 32.0% O
(b) Saran; 24.8% C, 2.0% H, 73.1% Cl
(c) polyethylene; 86% C, 14% H
(d) polystyrene; 92.3% C, 7.7% H
(e) Orlon; 67.9% C, 5.70% H, 26.4% N

5. A compound contains 4.07% hydrogen, 24.27% carbon, and 71.65 chlorine. Its molar mass is 98.96g. What are its empirical and molecular formula?

6. Find out the empirical formula and molecular formula in organic compounds that contain 62.07 % Carbon,10% Hydrogen, and 14% Nitrogen. The molecular mass of the compound is 114 gm/mole.

7. An unknown compound is found to contain 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen with a molecular mass of 60.0 g/mol. What is the molecular formula of the unknown compound?

8. A hydrocarbon is a compound comprised of carbon and hydrogen atoms. An unknown hydrocarbon is found to contain 85.7% carbon and an atomic mass of 84.0 g/mol. What is its molecular formula?

9. A piece of iron ore is found to contain a compound containing 72.3% iron and 27.7% oxygen with a molecular mass of 231.4 g/mol. What is the molecular formula of the compound?

10. A compound containing 40.0% carbon, 5.7% hydrogen, and 53.3% oxygen has an atomic mass of 175 g/mol. What is the molecular formula?

11. A compound contains 87.4% nitrogen and 12.6% hydrogen. If the molecular mass of the compound is 32.05 g/mol, what is the molecular formula?

LECT-9

STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

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 the products 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 :

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

Here, methane and dioxygen are called *reactants, and carbon dioxide and water are* called *products. All the reactants and* the products are gases in the above reaction and this has been indicated by a letter (g) in the brackets next to its formula. Similarly, in the case of solids and liquids, (s) and (l) are written respectively.

The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients. Similarly, the coefficient for CH_4 and CO_2 is one in each case. They represent the number of molecules (and moles as well) taking part in the reaction or formed in the reaction. Thus, according to the above chemical reaction,

- One **mole** of $CH_4(g)$ reacts with two **moles** of $O_2(g)$ to give one **mole** of $CO_2(g)$ and two **moles** of $H_2O(g)$
- One **molecule** of $CH_4(g)$ reacts with 2 **molecules** of $O_2(g)$ to give one **molecule** of $CO_2(g)$ and 2 molecules of $H_2O(g)$
- 22.7 L of $CH_4(g)$ reacts with 45.4 L of O_2 (g) to give 22.7 L of CO_2 (g) and 45.4 L of $H_2O(g)$
- 16 g of CH_4 (g) reacts with 2×32 g of O_2 (g) to give 44 g of CO_2 (g) and 2×18 g of H_2O (g).

From these relationships, the given data can be interconverted as follows :

From these relationships, the given data can be interconverted as follows :

$$\text{mass} \rightleftharpoons \text{moles} \rightleftharpoons \text{no. of molecules}$$

$$\frac{\text{Mass}}{\text{Volume}} = \text{Density}$$

LECT-10

Limiting Reagent

Many times, the reactions are carried out when the reactants are not present in the amounts required by a balanced chemical reaction. In such situations, one reactant is in excess over the other. The reactant which is present in the lesser amount gets consumed after some time and after that, no further reaction takes place whatever be the amount of the other reactant present. Hence, the reactant which gets consumed limits the amount of product formed and is, therefore, called the limiting reagent.

Concentration:

A majority of reactions in the laboratories are carried out in solutions. Therefore, it is important to understand how the amount of substance is expressed when it is present in the form of a solution. The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways.

1. Mass percent or weight percent (w/w %)
2. Mole fraction
3. Molarity
4. Molality

1. Mass per cent

It is obtained by using the following relation:

$$\text{Mass per cent} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute.

Solution

$$\text{Mass per cent of A} = \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$$

$$= \frac{2\text{ g}}{2\text{ g of A} + 18\text{ g of water}} \times 100$$

$$= \frac{2\text{ g}}{20\text{ g}} \times 100$$

$= 10\ \%$

2. Mole Fraction

It is the ratio of number of moles of a particular component to the total number of moles of the solution. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively; then the mole fractions of A and B are given as

Mole fraction of A

$$= \frac{\text{No. of moles of A}}{\text{No. of moles of solution}}$$

$$= \frac{n_A}{n_A + n_B}$$

Mole fraction of B

$$= \frac{\text{No. of moles of B}}{\text{No. of moles of solution}}$$

$$= \frac{n_B}{n_A + n_B}$$

3. Molarity

It is the most widely used unit and is denoted by M. It is defined as the number of moles of the solute in 1 litre of the solution. Thus,

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

Suppose we have 1 M solution of a substance, say NaOH and we want to prepare a 0.2 M solution from it.

1 M NaOH means 1 mol of NaOH present in 1 litre of the solution. For 0.2 M solution we require 0.2 moles of NaOH in 1 litre solution.

Hence, we have to take 0.2 moles of NaOH and make the solution to 1 litre.

Now how much volume of concentrated (1M) NaOH solution be taken which contains 0.2 moles of NaOH can be calculated as follows:

If 1 mol is present in 1 L or 1000 mL then 0.2 mol is present in

$$\frac{1000 \text{ mL}}{1 \text{ mol}} \times 0.2 \text{ mol}$$

= 200 mL

Thus, 200 mL of 1M NaOH are taken and enough water is added to dilute it to make it 1 litre. In fact for such calculations, a general formula, $M_1 \times V_1 = M_2 \times V_2$ where M and *V are* molarity and volume respectively can be used. In this case, M_1 is equal to 0.2; $V_1 = 1000$ *ml* and, $M_2 = 1.0$; V_2 is to be calculated. Substituting the values in the formula:

$0.2 \text{ M} \times 1000 \text{ mL} = 1.0 \text{ M} \times V_2$

$$\therefore V_2 = \frac{0.2\text{M} \times 1000\text{mL}}{1.0\text{M}} = 200 \text{ mL}$$

Note that the number of moles of solute (NaOH) was 0.2 in 200 mL and *it has remained the same, i.e., 0.2 even after dilution (in 1000* ml) as we have changed just the amount of solvent (i.e. water) and have not done anything with respect to NaOH.

4. Molality

It is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by m.

$$\text{Thus, Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$