Chapter I :

Fundamental concepts

Chemistry is the science that deals with matter (structure and properties of matter and transformations from one form of matter to another).

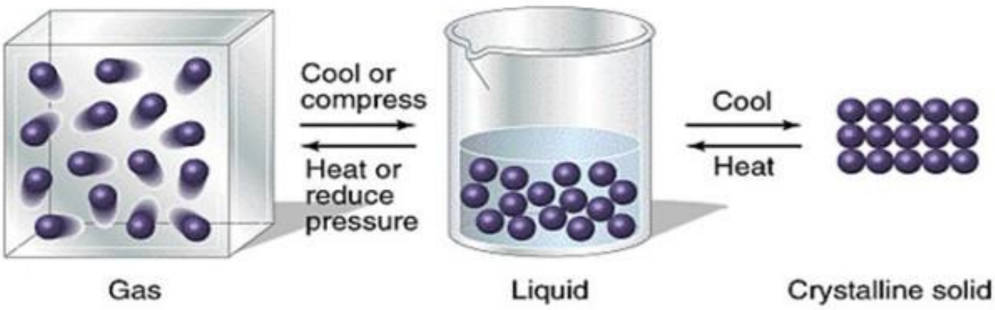
I-Definitions :

- **Matter:** is any thing that has mass and takes up space

States of Matter:

- **Gases:** have no definite shape or volume. They expand to fill whatever container they are put into. They are highly compressible.
- **Liquids:** have no definite shape, but they have a definite volume. They are slightly compressible.
- **Solids:** have definite shapes and definite volumes. They are incompressible.

Note: the state of a given sample of matter depends on the strength of the forces among the particles contained in the matter. Whether a substance is a gas, a liquid, or a solid depends on its temperature and pressure. We can convert state of matter from one form to another by changing temperature and pressure (however, the chemical identity of substance does not change – physical change).



- **Atom :** Atoms are the basic building blocks of matter. Anything that takes up space and anything with mass is made up of atoms.

- **Molecule :** the simplest unit of a chemical substance, usually a group of two or more atoms.
- **Mole (mol):** The amount of a substance that contains as many elementary particles (atoms,moleculesorions), where each mole has number of 6.022×10^{23}particles.

1 mole = 6.022×10^{23}particles =Avogadro's number N_a

1 mol Al = 6.02×10^{23} atoms

1 mol CO_2 = 6.02×10^{23} molecules

1 mol NaCl= 6.02×10^{23} Na^+ ions = 6.02×10^{23} Cl^- ions

- **Avogadro's number ($6.022\times \times\times \times10^{23}$):** number of formula units in a mole.

1 mole of hydrogen atoms = 6.022×10^{23} atoms of hydrogen

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

molecules of water 1 mole of Na^+ ions = 6.022×10^{23} ions of Na^+

- **Formula and molecular weight:** formula weight (FW) of a compound is the sum of the atomic weights in atomic mass units (amu) of all atoms in the compound's formula (for both ionic and covalent compounds). The molecular weight (MW) is the same as the formula weight; however, it is only used for the covalent compounds.

MW of H_2SO_4: 2×(1 amu) for H + 1×(32 amu) for S + 4×(16 amu) for O = 98 amu

MW or FW of $AlCl_3$: 1×(27 amu) for Al + 3×(35.5 amu) for Cl = 133.5 amu

- **Molar mass:** is the mass of one mole of the substance expressed in grams. We can say that it is the formula weight of a compound expressed in grams.

Formula weight of H_2O = 18 amu →molar mass = 18 g (mass of 1 mole H2O)

Formula weight of NaCl = 58.5 amu →molar mass = 58.5 g (mass of 1 mole NaCl)

- **Atomic weight (Atomic mass):** The atomic mass of an element is the average mass of the atoms of an element measured in *atomic mass unit* (amu, also known as *daltons*, D). The atomic mass is a weighted average of all of the isotopes of that element, in which the mass of each isotope is multiplied by the abundance of that particular isotope. (Atomic mass is also referred to as *atomic weight*, but the term "mass" is more accurate.)

- **Atomic mass unit (amu):** Atomic mass is expressed as a multiple of one-twelfth the mass of the carbon-12 atom, $1.992646547 \times 10^{-23}$ gram, which is assigned an atomic mass of 12 units. In this scale, 1 atomic mass unit (amu) corresponds to $1.660539040 \times 10^{-24}$ gram. The atomic mass unit is also called the dalton (Da), after English chemist John Dalton.

- **molar volume :** At standard Temperature and Pressure (STP) the molar volume (V_m) is the volume occupied by one mole of a chemical element or a chemical compound.

- **Lavoisier's Law of Conservation of Mass :** With the development of more precise ideas on elements, compounds and mixtures, scientists began to investigate how and why substances react. French chemist ***Antoine Lavoisier*** laid the foundation to the scientific investigation of matter by describing that substances react by following certain laws. These laws are called the laws of chemical combination. These eventually formed the basis of Dalton's Atomic Theory of Matter.

Law of conservation of mass: According to this law, during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants.

$$\underset{100\,g}{\text{Mercuric Oxide}} \longrightarrow \underset{92.6\,g}{\text{Mercury}} + \underset{7.4\,g}{\text{Oxygen}}$$

The law of conservation of mass is also known as the "law of indestructibility of matter."

- **Chemical Reactions :**

$$\text{Reactants} \rightarrow \text{Products}$$

Chemical reactions occur when chemical bonds between atoms are formed or broken. The substances that go into a chemical reaction are called the reactants, and the substances produced at the end of the reaction are known as the products.

$$2HgO\,(s) \rightarrow 2Hg\,(l) + O2(g)$$

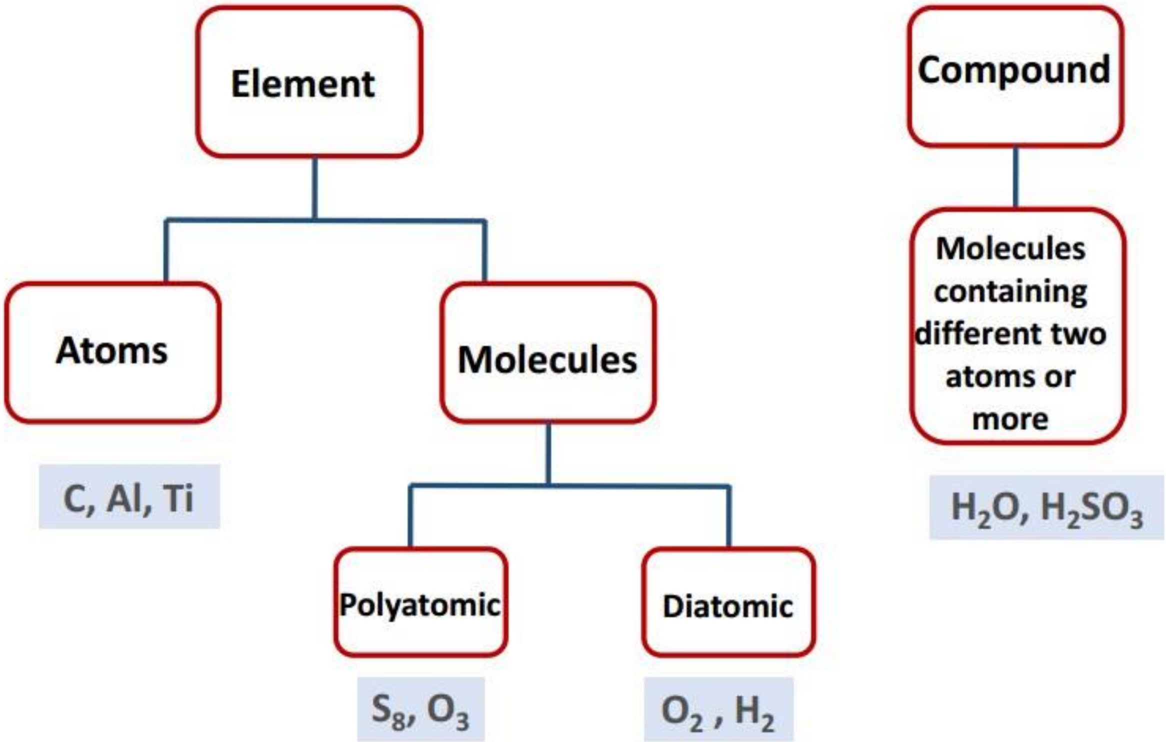
II- Qualitative aspect of the matter :

- **Element:** is a substance that consists of identical atoms (hydrogen, oxygen, and Iron). An element cannot be divided by chemical and physical methods. 116 elements are known (88 occur in nature and chemist have made the others inthe lab).

Symbols of elements: often an element's name is derived from a Greek, Latin, or German word that describes some property of the element. These symbols usually consist of the first letter or the first two letters of the elements name (F for fluorine and Ne for neon). Sometimes, however the two letters used are not thefirst two letters in the name (Zn for zinc).

- **Pure Substance:** a pure substance will always have the same composition and it cannot be more purified. Pure substances are either elements or compounds:

- **Compound:** is a pure substance made up of two or more elements in a fixed ratio by mass (for example, water H_2O and Sodium chloride NaCl).



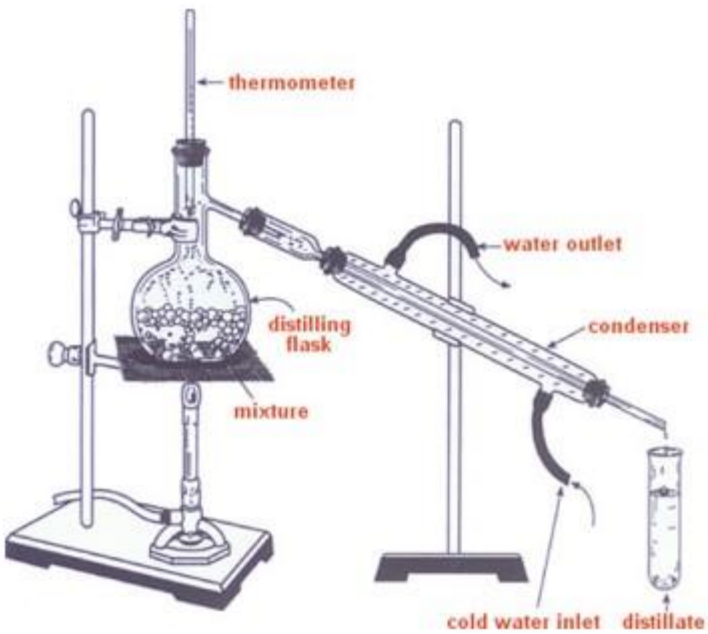
- **Mixture:** is a combination of two or more pure substances. Mixtures are divided into two groups:

1- Homogeneous: the mixture is uniform and throughout and no amount of magnification will reveal the presence of differentsubstances (for example, air, and salt in water).

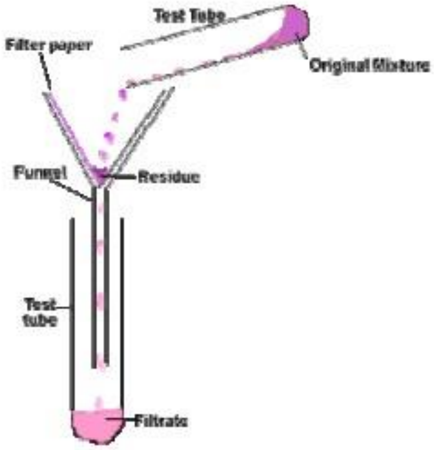
2- Heterogeneous: the mixture is not uniform. Therefore, it contains regions that have different properties from those of other regions (for example, soup, milk, and blood).

Note: if we know the physical properties of individual substances in a heterogeneous mixture, we can use appropriate physical means to separate the mixture into its component parts. Two methods of separation are distillation and filtration:

Distillation: Distillation is a widely used method for separating mixtures based on differences in boiling points of components of the mixture. To separate a mixture of liquids, the liquid can be heated to force components, which have different boiling points, into the gas phase. The gas is then condensed back into liquid form andcollected. Repeating the process on the collected liquid to improve the purity of the product is called double distillation.



Filtration: is commonly the mechanical or physical operation which is used for the separation of solids from fluids (liquids or gases) by interposing a medium through which only the fluid can pass. Oversize solids in the fluid are retained, but the separation is not complete; solids will be contaminated with some fluid and filtrate will contain fine particles (depending on the pore size and filter thickness).

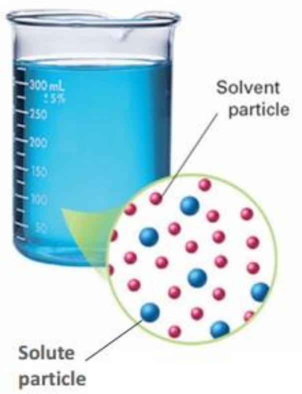


- **Solutions :**

Solution: a homogeneous mixture of two or more substances

Solute: a substance that is being dissolved (smaller amount)

Solvent: a substance which dissolves a solute (larger amount)



- **Saturated solution :**

A saturated solution is a chemical solution containing the maximum concentration of a solute dissolved in the solvent. The additional solute will not dissolve in a saturated solution.

- **Diluted Solution :**

Dilution is the process of decreasing the concentration of a solute in a solution, usually simply by mixing with more solvent like adding more water to the solution. To dilute a solution means to add more solvent without the addition of more solute.

III- Quantitative aspect of matter :

Concentrations :

The concentration of a solution is the amount of solute present in a given quantity of a solvent or solution.

- **Molarity :** The number of moles of solute dissolved In one liter of solution.

$$\boxed{\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}}$$

Example : A solution has a volume of 2.0 L and contains 36.0 g of glucose (C6H12O6). If the molar mass of glucose is 180 g/mol, what is the molarity of the solution?

No. of mol of glucose = wt (g) / Mw (g/mol) = 36.0 g/ 180g/mol= 0.2 mol

M = n (mol)/ V (L) = 0.2 mol/2.0 L= 0.1 mol/L

- **Molality :** The number of moles of solute dissolved in one kilogram of solvent

Molality *(m)*

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$$

Molarity *(M)*

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Example : What is the molality of a 5.86 Methanol (C_2H_5OH) solution whose density is 0.927 g/mL?

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$$

Assume 1 L of solution:

5.86 moles ethanol = 270 g ethanol

927 g of solution (1000 mL x 0.927 g/mL)

mass of solvent = mass of solution –mass of solute

= 927 g –270 g = 657 g = 0.657 kg

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{5.86 \text{ moles } C_2H_5OH}{0.657 \text{ kg solvent}} = 8.92\ m$$

For instance, if you have a solution that comprises 10 grams of salt dissolved in 100 grams of water, the mass per cent of salt in the solution would be:

→ **Mass percent = (10 g ÷ 110 g) × 100% = 9.09%**

This means that the solution contains 9.09% salt by mass.

- **Normality :** A solution is a homogenous mixture of two or more components. Each component of a homogenous solution has the same chemical composition. Therefore, the number of grams equivalent of an element dissolved in one litre of solution is the solution's normality.

Define Normality and Its Formula :

The term "normality" refers to the number of gram equivalents present in a litre solution. It is utilised in the preparation of acid or basic solutions. The units of concentration are denoted by the symbols N, eq/L.

$$N = \frac{Gram\ Equivalent\ of\ Solute}{volume\ of\ solution\ in\ litre}$$

$$N = \frac{Weight}{Equivalent\ weight} \times \frac{1000}{V\ (ml)}$$

$$Equivalent\ weight = \frac{\text{Molecular weight}}{X}$$

Number of Gram Equivalent : let's get forward with understanding the concept of gram equivalence.

We know that : **no of moles = Mass/Molecular weight**

Number of gram equivalent = Mass/Equivalent weight

Equivalent weight = Molecular weight / X (X = valence factor, where valence factor for acids and bases is the number of H^+ and OH^- ions they release in the solution, respectively).

We'll understand these two formulas with an example.

Let's find out Gram Equivalent.

Example 1 : Find the number of gram equivalents present in 0.5 g of HCl.

HCl releases one H^+ ion in the solution, so X = 1.

The molecular weight of HCl = 36.46 g.

So, **equivalent weight =Molecular weight /X** = 36.46/1 = 36.46 g,

and Number of gram equivalent= Mass/Equivalent weight = 0.5/36.46 = 0.0137

So, we get the **number of gram equivalent = 0.0137**

Example 2: Let us take another example of 1.06 g of Na_2CO_3 to understand this concept clearly

We are given the mass of Na_2CO_3 = 1.06 g.

Firstly, Find the equivalent weight of Na_2CO_3.

Since Na_2CO_3 is a **salt**, so the number of positive charges on the cation gives X = 2 Molecular weight = 106 g So, **Equivalent Weight = Molecular weight/X**= 106/2 = 53 g,

and **number of gram equivalent = Mass/Equivalent weight**= 1.06/53 = 0.02

In a chemical equation, the number of grams equivalent of both reactions always remains the same.

The following simple relationship exists between normality, *N*, and molarity, M,

$$N = n \times M$$

M is the molecular weight of moles per litre, and n is some equivalents formed. The number of equivalents is an integer for acid-base reactions, but maybe a fraction in a redox reaction.

- **Molar fraction :**It is a unit of concentration, defined to be equal to the number of moles of a component divided by the total number of moles of a solution. Because it is a ratio, mole fraction is a unitless expression. The mole fraction of all components of a solution, when added together, is equal 1.

$$X_i = \frac{n_i}{n_T}, \qquad \sum X_i = 1$$

- **Mass concentration :** ρ_i is defined as the mass of a constituent m_i divided by the volume of the mixture V. For a pure chemical the mass concentration equals its density (mass divided by volume); thus the mass concentration of a component in a mixture can be called the density of a component in a mixture. This explains the usage of ρ (the lower case Greek letter rho), the symbol most often used for density.

- **Mass Percentage Formula :** Mass Percent Fomula can be defined as the way that is used to express a concentration. Mass percent is a way to express the concentration of a substance in a mixture. It can further be defined as the mass of a solute in a mixture divided by the total mass of the mixture, and multiplied by 100%:

$$\text{Mass Percentage} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100$$