Chapter 1: General Concepts

1. The concept of matter

Matter is everything in the universe with a definite mass and occupies a specific space in the vacuum, such as water, trees, air…

Matter is further classified depending the nature and number of substances making them up. They can be classified in to two broad categories: Pure substances and Mixtures. These can be further sub-divided as shown in Fig. 1.

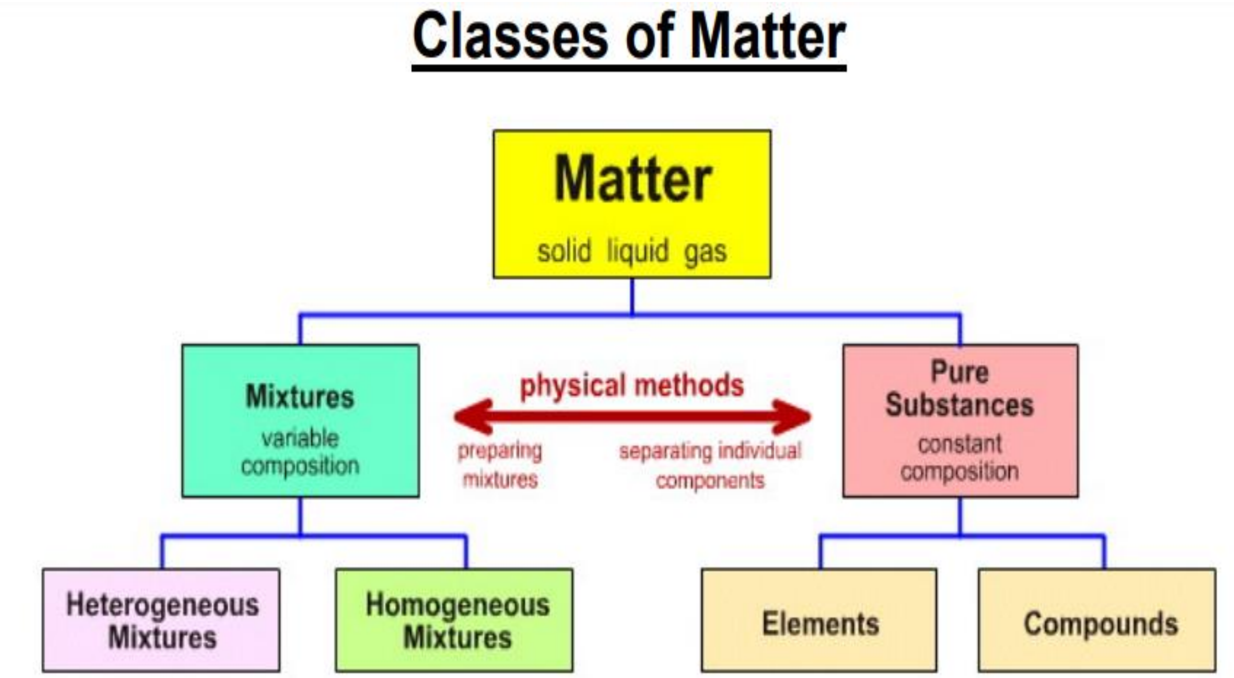


Fig. 1

1.1.Mixtures

When we mix two or more substances (elements or compounds) together, we obtain a mixture with a chemical composition and physical properties (mass, volume, density, boiling point) that vary depending on the sample studied. Mixtures can be classified into two types:

a) Homogeneous Mixtures: A homogeneous mixture is a mixture in which its components can't be distinguished by the naked eye. Examples include a mixture of gases, seawater, and mineral water.

b) Heterogeneous Mixtures: A heterogeneous mixture is a mixture in which its components cannot be distinguished by the naked eye. Examples include oil and water, fog, and vinegar mixed with oil

1.2. Pure substances

A pure substance is a material composed of only one type of chemical component, which means it is homogeneous in its composition. A pure substance can either be a single element or a compound.

a) Elements: Particles of an element consist of only one type of atoms. These particles may exist as atoms or molecules. For example: Carbon, sodium, copper, silver, hydrogen, oxygen, etc. are elements. Their all atoms are of one type Example. Oxygen-O_2. They cannot be broken down into two or more elements.

b) Compounds: It consists of two or more types of atoms combined together to form a compound and often breaks down into simpler substances. Examples include HCl and NH_3.

2. Concept of Molecule and Atom

2.1. Molecule

A molecule is the smallest part of a pure substance that retains the characteristics of the substance, including its composition, physical properties, and chemical properties.

2.2. Atom

An atom is the smallest part of a molecule, and it has very small dimensions measured in angstroms (Å).

$$1 \text{Å} = 10^{-10} \text{ meters} = 10^{-8} \text{ centimeters.}$$

3. The concept of the mole

Is a unit of measurement for the quantity of a substance. One mole of a substance is equivalent to the gram molecular weight or the gram atomic weight. One mole of any substance contains Avogadro's number ($N_A = 6.023 \times 10^{23}$) of constituent particles, whether these particles are atoms, molecules, ions, or electrons.

For example:

- 1 mole of oxygen atoms contains 6.023×10^{23} oxygen atoms and weighs 16 grams.
- 1 mole of water molecules (H_2O) contains 6.023×10^{23} water molecules and weighs 18 grams.

To calculate the number of moles of a substance, we use one of the two relationships:

Number of moles = Mass of substance / Molar mass of substance

$$n\,(mol) = \frac{m\,(g)}{M\,(g/mol)}$$

Number of moles = Number of particles / Avogadro's number.

$$n\,(mol) = \frac{N}{N_A}$$

4. Atomic mass (relative atomic mass)

Is the mass of a single atom of an element, expressed in atomic mass units (amu or u, sometimes written without units). It represents the sum of the protons and neutrons present in the atom. Electrons contribute a very small mass and are generally not considered in the calculation.

The atomic mass unit (uma) is defined by the following relationship:

1 uma = 1/12 * mass of one ${}_{6}C^{12}$ atom (g)

Find the mass of a carbon atom ${}_{6}C^{12}$ in grams:

$$1 \text{ mole of C} \longrightarrow N_A \text{ of C atoms} \longrightarrow 12\text{ g}$$

$$1 \text{ atom of C} \longrightarrow m_{atom}\text{ (g)}$$

Hence, $m_{atom} = \frac{12}{N_A} g$

$$1\ amu = \frac{1}{12} \times \frac{12}{N_A} = \frac{1}{N_A} \Rightarrow 1\ amu = 1.66 \times 10^{-24} g = 1.66 \times 10^{-27} Kg$$

Calculating the mass of a carbon atom (m_C) in amu:

$$1\ amu \rightarrow \frac{1}{N_A}\ g$$

$$m_c\ amu \rightarrow \frac{12}{N_A}\ g$$

$$\Longrightarrow m_c = \frac{12}{N_A} \times N_A = 12\ amu$$

In general:

$$\text{The atomic mass of an element in amu} = \frac{\text{Mass of an atom of that element}}{\frac{1}{12} * \text{mass of one } {}_{6}C^{12} \text{ atom (g)}}$$

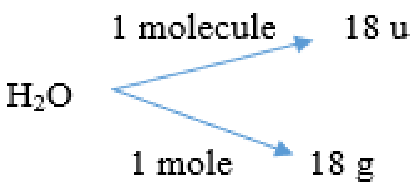
Note:

1 mole of C $\longrightarrow$ 12 g

The mass of a single atom of C $\longrightarrow$12 amu

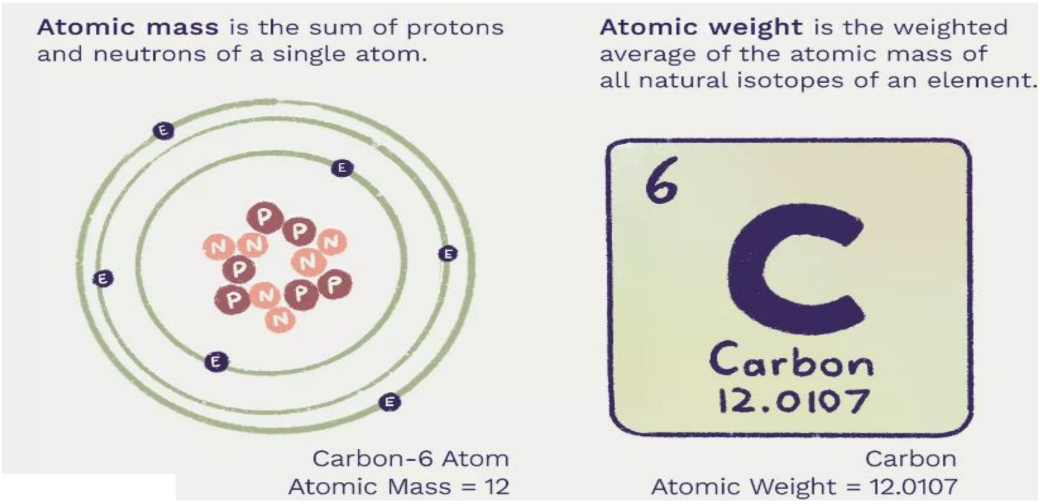
It is observed that the mass of an atom (or molecule) measured in amu numerically equals the mass of 1 mole of the same atom (or molecule) measured in grams.

Example:



4.1.The atomic weight (relative atomic mass)

Atoms of the same chemical element may contain different numbers of neutrons, resulting in different atomic masses. Therefore, the atomic weight of an element is the average of the atomic masses of its natural isotopes.



4.2. Gram Atomic Mass

The mass of 1 mole of atoms of an element in grams.

4.3. Molecular Mass

It is the sum of the atomic masses of the constituent atoms of a molecule.

4.4. Gram Molecular Mass

The mass of 1 mole of molecules of a compound in grams.

Example:

Let it be 1.68 g of iron (Fe) (M = 56 g). Find the number of moles, the number of atoms, and the atomic mass of an iron atom in atomic mass units (u).

5. Fundamental Laws

5.1. Laws of weight

The law of mass conservation (Lavoisier's law - Lavoisier 1774)

During a chemical reaction, there is neither a loss nor a gain of mass. In other words, mass is conserved, and it is not created or destroyed; rather, it can be transformed from one form to another. This means that the total mass of the reacting substances equals the total mass of the resulting substances. This law is commonly used in writing chemical equations, taking into consideration the equal number of atoms of each element on both sides of the equation.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

$$\begin{array}{ccc} C=1 & & C=1 \\ H=4 & = & H=4 \\ O=4 & & O=4 \end{array}$$

The Law of Definite Proportions (Proust's law 1799)

The Law of Definite Proportions, also known as Proust's law, states that in a chemical compound, the elements are always present in fixed and definite proportions by mass. This means that regardless of the source or method of preparation, a compound will always contain the same elements in the same proportion by mass.

The mass percentage can be calculated using the weight of the element and the weight of the

$$\%\ \text{mass of element} = \frac{mass\ of\ element}{total\ mass\ of\ compound} \times 100$$

compound in grams, according to the following formula:

Example:

Pure water, whether obtained from groundwater, rainwater, or chemical reactions, always contains 11.1% of its mass as hydrogen and 88.90% as oxygen. This means that the ratio of one oxygen atom to two hydrogen atoms (H_2O) remains constant in water regardless of its source or method of production.

The Law of Multiple Proportions (Dalton's law - 1804)

States that when two elements combine to form multiple different compounds, the ratio between the masses of one element that combines with a fixed mass of the other element in these compounds is a simple and whole-number ratio. This ratio can be expressed using numbers separated by a colon (such as 2:3) or in fractional form.

Law of Multiple Proportions = Mass ratio of an element in one compound / Mass ratio of the same element in another compound = Whole number

Mass ratio of element A in the compound = Mass of element A in the compound / Mass of element B in the compound

Example 1:

Hydrogen can combine with oxygen to form one of the following compounds:
water (H_2O) or hydrogen peroxide (H_2O_2).
For instance, 2 g of hydrogen can combine with 16 g of oxygen to form water, and 2 g of hydrogen can combine with 32 g of oxygen to form hydrogen peroxide.
Thus, the ratio of the oxygen mass in the first compound to the second compound is 16:32, which simplifies to 1:2. This ratio is a simple and whole-number ratio.

Example 2:

The compound	Cu%	Cl%	The mass of Cu(g) in 100 g of the compound	The mass of Cl(g) in 100 g of the compound	The mass ratio(m_{Cu}/m_{Cl})
I	64.20	35.80	64.20	35.80	1.793
II	47.27	52.73	47.27	52.73	0.8964

$$\frac{\textit{The mass ratio of compound I}}{\textit{The mass ratio of compound II}} = \frac{1.739}{0.8964} = 2$$

The mass ratio of copper in the two compounds 2:1

The Gas Laws - The Ideal Gas Equation (Clapeyron,Equation)

If we have n moles of an ideal gas taken at temperature T and in volume V, it is influenced by pressure P, where :

$$P.V = n.R.T$$

- P: Gas pressure
- V: Gas volume
- n: Number of moles of the gas
- R: Universal gas constant for ideal gases
- T: Absolute temperature

Boyle's Law (Boyle - Mariotte Law)

The volume of a fixed amount of gas at a constant temperature is inversely proportional to the pressure.

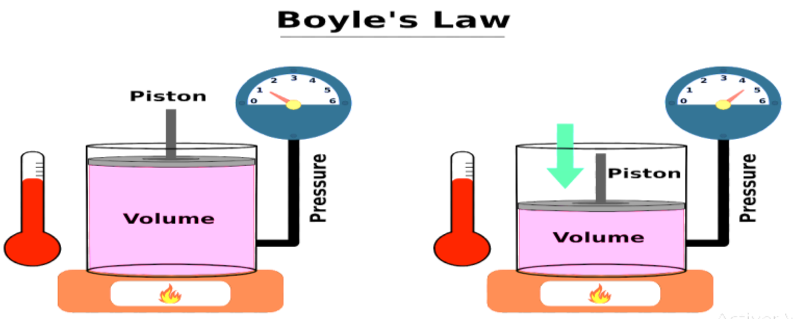
If we have the following conditions:

Initial conditions: $P_1.V_1 = n.R.T$

Final conditions: $P_2.V_2 = n.R.T$

Then, we have $P_1.V_1 = P_2.V_2 = K$ / the constant K:

This law describes the relationship between the volume and pressure of a gas at constant temperature, asserting that the product of pressure and volume for a given amount of gas remains constant.



Charles's Law

Charles formulated his law as follows: "The volume of a fixed amount of gas is directly proportional to the absolute temperature at constant pressure."

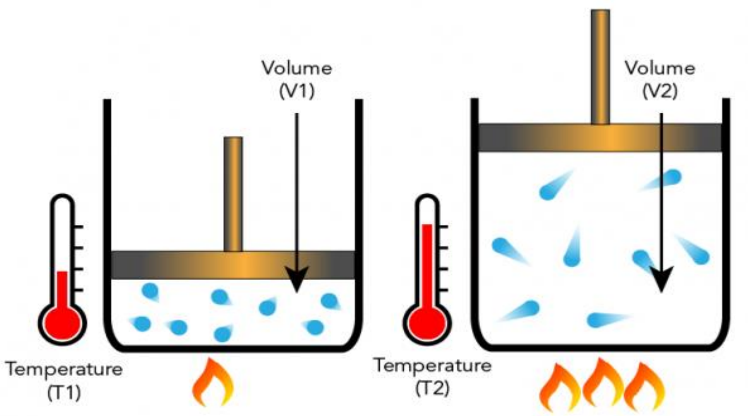
If we have the following conditions:

Initial conditions: $P.V_1 = n.R.T_1$

Final conditions: $P.V_2 = n.R.T_2$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad \frac{V_2}{V_1} = \frac{T_2}{T_1} \quad \text{or} \quad V_1 \cdot T_2 = V_2 \cdot T_1$$

The image below shows how adding heat makes molecules move faster and hit the sides and lid with greater force, thus moving the lid up as the gas expands.



Gay-Lussac's Law

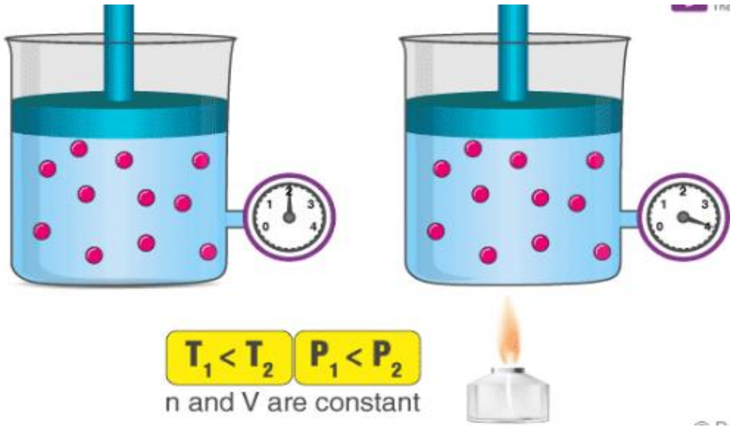
Gay-Lussac's law states: "The pressure of a fixed amount of gas is directly proportional to the absolute temperature at constant volume."

If we have the following conditions:

Initial conditions: $P_1.V = n.R.T_1$

Final conditions: $P_2.V = n.R.T_2$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{or} \quad \frac{P_1}{P_2} = \frac{T_1}{T_2}$$



Dalton's Law of Partial Pressures

This law states that the total pressure of a mixture of non-reacting gases is equal to the sum of the partial pressures of all the gases in the mixture.

$P = P_1 + P_2 + P_3 + \ldots$

Where P_1, P_2, P_3 are partial pressures.

6. Gas Density (It is referred to as Relative Density in English sources)

The density of gaseous substances is measured relative to air.

$$d_{\text{gas}} = \frac{\rho_{gas}}{\rho_{air}}$$

In standard conditions, the Volumic Mass (The specific volume or volume-specific mass in English references) of air (ρ_{air}) is 1.29 g/L.

If the quantity of substance for the gas is n=1 mole.

$$\rho_{gas} = \frac{M_{gas}}{V_{Mgas}}$$

$$d_{gas} = \frac{M_{gas}}{1.29.V_{Mgas}}$$

Under standard conditions, V_M=22.4 *l/mol* by substitution we find:

$$d_{gas} = \frac{M_{gas}}{29}$$

Similarly, under standard conditions, the mass of one mole of air $M_{air} = \rho_{air}.V_{M/air} = 29g/mol$

$$d_{gas} = \frac{M_{gas}}{29} = \frac{M_{gas}}{M_{air}}$$

7. The solution

A solution is a homogeneous mixture composed of a solvent and a solute. There are methods through which one can express the quantity of the solvent and solute and the relationship between them.

7.1. Methods for expressing the concentration of a solution

Molarity (Molar concentration or molar concentration)

Molarity is the number of moles of solute present in one liter of the solution. It is represented by "M" or "C." Its unit is mol/L or M (molar).

$$(M) = \frac{n_2}{V_{sol}(L)}$$

Molality (Molal Concentration)

Molality, represented by "m," is the number of moles of solute present in 1 kilogram of solvent. Its unit is mol/kg or molal.

$$m = \frac{n_2}{m_{1(Kg)}}$$

Mass Concentration

Mass concentration, represents the mass of the solute present in 1 liter of the solution. It is denoted by "C_m" and its unit is g/L (grams per liter).

$$\underbrace{C_m}_{g.L^{-1}} = \frac{m \ (g)}{V \ (L)}$$

Mass Percent of the Solute

The mass percentage of the solute is the mass of the Solute measured in grams present in 100g of the solution.

$$\text{Mass percent} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution(g)}} \times 100\%$$

The mass fraction

Is the ratio of the mass of a component in the mixture to the total mass of the mixture.

$$w_i = \frac{m_i}{m_T} = \frac{m_i}{\sum m_i}$$

- For a solution, it is the ratio of the mass of a solute to the mass of the solution.

$w_{solute} = \frac{m_{solute}}{m_{solvent}+m_{solute}}$ and

$$w_{solvent} = \frac{m_{solvent}}{m_{solvent} + m_{solute}}$$

The sum of the mass fraction of each component is always equal to one.

$$\sum w_i = 1$$

The molar fraction of the solute

The mole fraction of a substance in the mixture is the ratio of the mole of the substance in the mixture to the total mole of the mixture.

$$x_i = \frac{n_i}{\sum n_i}$$

For a solution, it is the ratio between the number of moles of the solute and the total number of moles of the solution.

$$x_{solute} = \frac{n_{solute}}{n_{solvent} + n_{solute}}$$

Normality" (Equivalent Concentration - Molar Concentration)

Is the number of gram equivalents of the solute present in 1 liter of the solution. Its unit is Eq/L or N.

$$N = \frac{Eq_2}{V_{sol}(L)}$$

Eq_2 is the number of gram equivalents of the Solute.

You can calculate the number of gram equivalents (Eq_2) by dividing the Weight of the solute in grams by its equivalent weight using the following relationship:

$$Eq_2 = \frac{Weight\ of\ Solute\ in\ grams}{Equivalent\ \mathrm{Weight}} \left[\frac{g}{g/Eq} \equiv Eq\right]$$

Where the equivalent weight of the solute can be calculated using the following relationship:

$$Equivalent\ Weight = \frac{Molar\ mass\ of\ the\ solute}{Valency} \left[\frac{g/mol}{Eq/mol} \equiv g/Eq\right]$$

Valency: It is the number of H^+ or OH^- ions exchanged or the number of electrons involved in oxidation-reduction reactions.

Example

$$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$$

$$Equivalent\ Weight = \frac{M_{H2SO4}}{2} = \frac{98}{2} = 49\ \mathrm{g/Eq}$$

The relationship between molarity (C) and normality (N)

$$N = \frac{Eq_2}{V_{sol}} = \frac{m_{solute}}{M_{solute} \times V} \times Valency = \frac{n_{solute}}{V} \times Valency = C \times Valency$$

The relationship between the number of gram equivalents and the number of moles

$$Eq_2 = \frac{m_{solute}}{M_{solute}} \times Valency = n_{solute} \times Valency$$

Density of Solution

If the sample under study is solid or liquid, the reference is water.

$$d_{sample} = \frac{\rho_{sample}}{\rho_{water}}$$

Note:

Density equals Volumic Mass if the latter is expressed in g/cm^3.

Examples:

1- A sample of an ideal gas with a volume of 5 L is under a pressure of 15 atm. Calculate the volume of this gas if its pressure becomes 3 atm, assuming constant temperature. By applying Boyle's law.

2- How many grams of sodium hydroxide (NaOH) are present in 500 mL of a solution with a concentration of 0.0412?