Chapter One: fundamental concepts

Definition of matter: Matter makes up all substances that have a mass and occupy space.

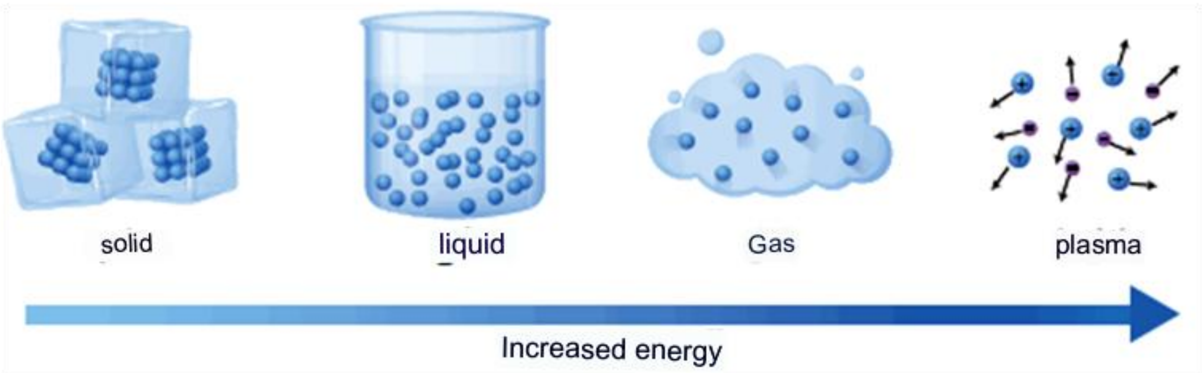
States of matter: A substance can exist in different physical forms, depending on its properties, temperature and pressure. These forms are called states of matter. The three best known are solids, liquids and gases. However, there are others. We can think of plasma which is almost a gas, with the difference that it contains free electrons and ions (atoms having lost electrons).

A- Solid state: It is a substance whose shape and size are fixed, meaning that the molecules that make up this substance are semi-fixed, close together, and arranged according to a specific system (for example, table salt).	 A solid has a definite shape and a definite volume
B- Liquid state: A liquid is a substance that takes the shape of the container in which it is placed while the volume remains constant. The molecules in a liquid body can move at a limited rate (for example, water).	 A liquid has an indefinite shape, it takes the shape of its container and a definite volume

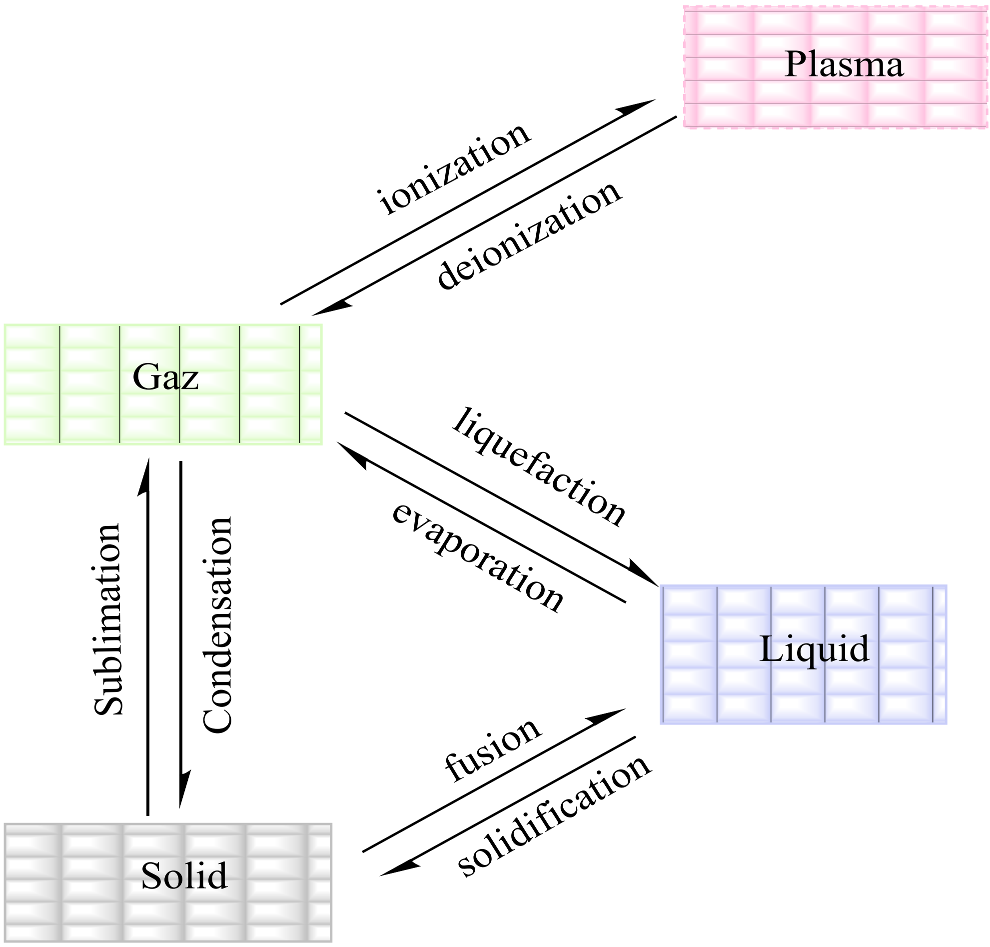
C- The gaseous state: The substance in this state is characterized by the property of expansion, as the gas takes the shape and size of the container in which it is placed, meaning that the molecules that make up this substance are completely free among themselves and are in chaotic movement (for example, oxygen gas	 A gas has an indefinite shape and an indefinite volume, it assumes the shape and volume of its container

Changes in states of matter:

Depending on the external conditions (temperature and pressure), the same substance can be either a solid, a liquid or a gas. In fact, the passage of matter from the **solid** state to the **liquid** state occurs by **melting**, from the **liquid to** the **gaseous** state by **vaporization** and from the **solid to** the **gaseous** state by **sublimation**. When particles of matter gain energy, their vibrational movement increases, and the bonding forces between their particles weaken, so matter moves from one state to another.



- A diagram showing the changes between states of matter



III- Classification of matter

Matter exists in the form of mixtures (homogeneous or heterogeneous) and pure substances (simple or compound).

1- **Pure substances:** They are substances composed of identical molecules. Pure substances are distinguished by their physical properties (melting point, boiling point, density, refractive index, etc.) or chemical properties.

They are divided into two types:

- **Simple pure substances:** They are substances that consist of one type of atom (H_2, O_2, Fe, etc.).

- **Pure compound substances:** consisting of molecules with different atoms (H_2O, sodium chloride NaCl, carbon dioxide CO_2, etc...)

2- Mixtures: They are compounds whose molecules are different, meaning they consist of two or more pure substances. They are divided into two categories:

- **Heterogeneous mixture:** through which it is possible to distinguish, by eye or with the help of a microscope, parts with different properties, meaning that it consists of two or more phases.

Example: A mixture of water and oil.

- **Homogeneous mixture:** It is a mixture of two substances in one phase such that its components cannot be distinguished.

Example: Salt water (a solution of water and salt) is a homogeneous mixture (the salt dissolves in water, so we obtain one phase). , air, steel (alloy), etc

A summary of how to distinguish the different main classifications of matter is presented in the following figure.

Matter

Does it have constant properties and compositions?

No

Mixture

- Physical combination of two or more substances
- Composition can vary
- Properties can vary with composition

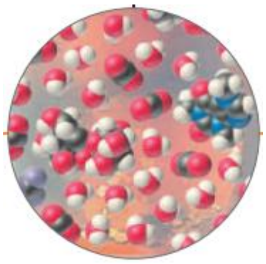
Is it uniform throughout ?

yes

Homogeneous

- One visible phase
- Same properties throughout

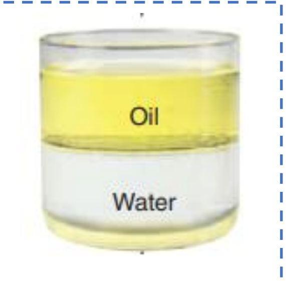
Soft drink

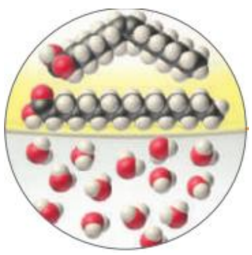


No

Heterogeneous

- Two or more visible phases
- Different properties in different phases





yes

Pure substances

- Only one substance present
- Definite and constant composition
- Properties always the same under the same conditions

Can it be simplified chemically ?

No

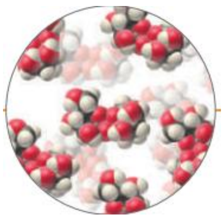
Elements

- Cannot be broken down into simpler substances by chemical or physical means
- Building blocks for all other types of matter
- 117 elements known

yes

Compound

- Can be broken down into constituent elements by chemical, but not physical, means
- Chemical combination of two or more elements Have definite, constant, elemental composition



Molecular representation

IV. Concept of atom, molecules, mole and Avogadro's number

IV. 1. The atom: A simple body consists of identical atoms. We can consider that an atom is the smallest particle in a simple body that we can imagine without destroying this simple body. There are more than a hundred atoms. These are based on the same model and differ in their physical and chemical properties. Each element or atom is represented by a symbol. The symbol consists of an uppercase letter or two letters, the first uppercase followed by a lowercase letter. This formalism must be strictly adhered to otherwise the elements are mixed together (e.g. Co = cobalt, while CO = carbon + oxygen).

Example: carbon (C), oxygen (O), copper (Cu), aluminum (Al). The size of an atom is very small in angstroms (1A°= 10^{-10}m) and its mass is about 10^{-26} kg

IV. 2. The molecule: Molecules are groups or clusters of atoms held together by means of chemical bonding. There are two types of molecule; molecules of an element and molecules of a compound

A molecule is a set of linked atoms. The molecule constitutes the smallest part of a pure body, Such as NaCl, O_3, H_2SO_4, etc..

IV. 3. The mole and Avogadro's number

3.1. Definition and unit: The mole designates the quantity of matter of a substance contained in a system which has as many elementary entities as there are atoms in 12g of carbon 12(^{12}C).

The mole is a unit of account used by chemists to express quantities of matter. Its unit is the mole (mol).

3. 2. Avogadro's number: It corresponds to the number of particles found in exactly 12 g of carbon 12. Also, the mole corresponds to the quantity of matter found in 12 g of carbon 12. Avogadro's number, symbolized N_A, corresponds to the number of particles found in a mole, i.e. 6.022×1023 particles.

To determine the number of moles from Avogadro's number and the number of particles, we can use the following formula:

$$n = \frac{N}{N_A} \quad \text{ou} \quad N = n \times N_A$$

Where

n: represent the number of moles (mol)

N: represent the number of particles (atoms, molecules, ions, etc.)

N_A: represent Avogadro's number

3. 3. Molar mass: The atomic molar mass of a chemical element is the mass of one mole of atom of this element taken in its natural state. It is expressed in g/mol or g.mol-1 and is denoted by the letter M. The different values of M are read in the periodic table. Example: M_H = 1g/mol, M_O = 16g/mol, M_{Ag} = 108g/mol.

3. 4. The molecular molar mass: or mass of a chemical substance is the mass of one mole of a molecule of that substance. It is determined by taking the sum of all the atomic molar masses of all the chemical elements present in the molecule. **Example:** Determine the molecular molar masses of: H_2O, $C_{22}H_{22}O_{11}$, $C_6H_5NO_2$, $CO(NH_2)_2$.

We give in g/mol: C = 12, H = 1, O = 16, N = 14.

3. 5. Atomic mass units (amu): An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12.

$$1uma = \frac{1}{12} m_{C^{12}}$$

Where $m_{C^{12}}$: represent the mass of one atom of carbon

$$\mathbf{12g \longrightarrow N_A\ atom\ of\ carbon}$$

$$\mathbf{m_{C^{12}}}\ (\ mass\ of\ one\ atom\ of\ carbon) \longrightarrow \mathbf{1 atom\ of\ carbon}$$

$$m_{C^{12}} = \frac{M_{C^{12}}}{N_A}$$

So
$$1mau = \frac{1}{N_A} = \frac{1}{6,023\,.\,10^{23}} = 1,666\,.\,10^{-24}g = 1,666\,.\,10^{-27}kg$$

Note: The key difference between atomic mass unit and atomic mass is that the atomic mass unit is the unit that we use to measure the mass of an atom whereas the atomic mass is the mass of a particular single atom.

3. 6. Molar volume (V_m):

It is the volume occupied by one mole of a substance (chemical element or chemical compound) in the gaseous state and under a certain temperature and pressure under regular conditions (P = 1atm). It is found that the volume equals 22.4 litres.

3. 7. Concept of quantity of matter n: The quantity of material n of a sample is obtained by dividing the mass m of this sample by the molar mass M of the chemical species considered.

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

Chemical reaction concept:

Definition: A chemical reaction is a chemical transformation during which pure substances disappear and simultaneously new pure substances are formed. The pure substances which disappear during a chemical reaction are called the reactants and the pure substances which are formed are called the products.

We can write an equation that shows the balance of a chemical reaction:

$$\overbrace{a\,A + bB}^{\text{Reactants}} \longrightarrow \overbrace{cC + dD}^{\text{Products formed}}$$

This balanced chemical equation obeys two laws:

- **Conservation of elements:** During a chemical reaction, there is conservation of atoms (the elements are conserved)

Example

$$S + O_2 \longrightarrow SO_2$$

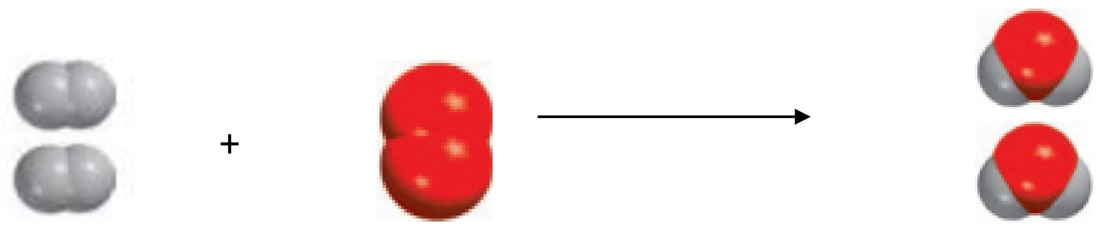
$$Fe + S \longrightarrow FeS$$

$$H2 + \frac{1}{2}O_2 \longrightarrow H_2O \quad \text{either} \quad 2H_2 + O_2 \longrightarrow 2H_2O$$

- **Mass conservation:** In a chemical reaction, the mass of the reactants disappeared is equal to the mass of the products trained (Lavoisier's law)

Example

2 H_2 react with 32.00 g of O_2 to give 36.04 g of H_2O



Two hydrogen molecules One oxygen molecule two water molecules

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Exercise :

Balance the following reactions:

$$Al + O_2 \longrightarrow Al_2O_3$$

$$Cu_2S + Cu_2O \longrightarrow Cu + SO_2$$

$$C + CO_2 \longrightarrow CO$$

$$H_2 + N_2 \longrightarrow NH_3$$

VI. Qualitative and quantitative aspect of the material

The main difference between qualitative and quantitative analysis chemistry is that qualitative analysis determines whether or not different chemical components are present in a sample, whereas quantitative analysis determines the amount of different chemical components present in a sample.

VI.1 Solutions

A solution is a homogeneous mixture of two or more components, where the all the components appear as a single phase.

Solvent:

The component that dissolves the other component is called the solvent.

Solute:

The component(s) that is/are dissolved in the solvent is/are called solute(s).

Generally solvent is present in major proportion compared to the solute. The amount of solute is lesser than the solvent.

Example of solution

- Sugar syrup is a solution where sugar is dissolved in water using heat. Here, water is the solvent and sugar is the solute.

VI. 2. Solution concentration units

Because solutions are mixtures, they have a variable composition. Specifying what the composition of a solution is involves specifying solute concentrations.

A concentration is the amount of solute present in a specified amount of solution. Many methods of expressing concentration exist, and certain methods are better suited for some purposes than others.

Molarity: The molarity of a solution is the number of moles of the solute in 1 litre of the solution.

$$molarity\ (C) = \frac{moles\ of\ solute)}{\text{litres of solution)}} = \frac{\text{n (mol)}}{\text{V(L)}}$$

$$n = \frac{mass\ (g)}{M}$$

Where: **MM = molar mass of solute**

The equation can also be used to find n (the number of moles of solute) if M (the molarity) and V (the volume in litres) are known, and V if n and M are known

$$n = MV \qquad V = \frac{n}{M}$$

Example 1:

A sample of $NaNO_3$ weighing 0.38 g is placed in a 50.0 mL volumetric flask. The flask is then filled with water to the mark on the neck, dissolving the solid. What is the molarity of the resulting solution?

Solution: To calculate the molarity, you need the moles of solute. Therefore, you first convert grams $NaNO_3$ to moles.

we have : $n = \frac{mass\ (g)}{MM}$

Where: $M = M_{Na} + M_N + 3\ M_O = 23 + 14 + 3 \times 16 = 85g/mol$

$$n = \frac{0.38}{85} = 4.47 \times 10^{-3} mol$$

So the molarity is: $C = \frac{\text{n (mol)}}{\text{V(L)}} = \frac{4.47 \times 10^{-3}}{50 \times 10^{-3}} = 0.089\ mol/l$

Example 2:

An experiment calls for the addition to a reaction vessel of 0.184 g of sodium hydroxide, NaOH, in aqueous solution. How many milliliters of 0.150 M NaOH should be added?

Solution:

Because molarity relates moles of solute to volume of solution, you first need to calculate the number of moles of NaOH. Then, you can calculate the volume in liter of solution, using the molarity equation.

$$n = \frac{mass\ (g)}{MM}$$

Where: $M_{NaOH} = M_{Na} + M_H + M_O = 23 + 1 + 16 = 40g/mol$

$$n = \frac{0.184}{40} = 4.6 \times 10^{-3} mol$$

So the volume is: $V = \frac{n}{M} \Longrightarrow V = \frac{4.6 \times 10^{-3}}{0.150} = 3.07 \times 10^{-2}\ L = 30.7ml$

exercice:

Exercice:

(a) Calculate the molarity of a solution of 0.25 mole of NaOH in 5.0 L of solution.

(b) Calculate the number of moles of citric acid in 250 mL of a 0.400 *M* solution of citricacid.

(c) Calculate the volume (in mL) of a 0.355 *M* NaOH solution which would contain 0.200 mole of NaOH. Calculate the mass of Na_2CO_3 that must be used to make 700 mL of a 0.136 *M* Na_2CO_3 solution.

(d) How many moles of sodium chloride should be put in a 50.0-mL volumetric flask to give a 0.15 M NaCl solution when the flask is filled to the mark with water? How many grams of NaCl is this?

Mass Percent Concentration: Percent by mass is the mass of solute in a solution divided by the total mass of solution, multiplied by 100.

$$\mathbf{m\ \% = \frac{m_{solute}}{m_{solvent}}\ x\ 100}$$

Example 1: A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$m\% = \frac{m_{NaCl}(g)}{m_{solution}(g)} = \frac{36.5}{355} \times 100\% = 10.3\%$$

Example 2: A dextrose (also called D-glucose, C6H12O6) solution with a mass of 2.00 × 102 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Normality (N): Represents the number of equivalents contained in one liter solution

$$\boldsymbol{Normality(N) = \frac{Number\ of\ equivalents\ of\ solute}{volume\ of\ solution(liter)}}$$

$$\mathbf{N} = \frac{\text{number of equivalent}}{\textbf{volume}} = \frac{\textbf{n eq g}}{\textbf{V}}$$

$$\boldsymbol{Number\ of\ gram\ equivalent = \frac{Mass}{Equivalent\ weight}}$$

$$\boldsymbol{Eq.mass = \frac{Molar.mass}{Valency}}$$

$$Equivalent\ weight = \frac{M}{z}$$

Z: is the valence factor, where valence factor for acids and bases is the number of H+ and OH- ions they release in the solution, respectively

Gram equivalent **(**Eq.mass or eq. weight)**:** It is the mass of matter corresponding to a mole of particles, whether they are H^+ protons or e- electrons. There are three families of gram equivalents classified according to fixation or loss of particles.

$$\boldsymbol{Eq.mass = \frac{Molar.mass}{Valency}}$$

$$Equivalent\ weight = \frac{M}{X}$$

A) - The gram equivalent of the acid:

$$\mathbf{eq\ g_{acide}} = \frac{\mathbf{M_{acide}}}{\mathbf{z_{H^+}}}$$

Example: Gram equivalent of hydrochloric acid HCl,

$$HCl \longrightarrow 1\,H^+ + Cl^-$$

$$Eq.g\ HCl = \frac{M_{HCl}}{nH^+} = \frac{36{,}5}{1} = 36{,}5\ g$$

B) Gram equivalent of base

$$\mathbf{Eg_{base}} = \frac{\mathbf{M_{base}}}{\mathbf{z_{OH-}}}$$

Example: Gram equivalent of calcium hydroxide $Ca(OH)_2$

$$Ca(OH)_2 \longrightarrow Ca^{2+} + 2\,OH^-$$

$$Eq.g\ Ca(OH)_2 = \frac{M_{Ca(OH)_2}}{n_{OH^-}} = \frac{74{,}1}{2} = 37{,}05\ g$$

c- Gram Equivalent in (oxidation – reduction) reaction (Redox)

$$\mathbf{Eq.g_{ox/red}} = \frac{\mathbf{M_{ox/red}}}{\mathbf{z_{e^-}}}$$

Example:

$$2KMnO_4 + 10FeSO_4 + 8H_2SO_4 \rightarrow 5Fe_2\,(SO_4)_3 + 2MnSO_4 + K_2SO_4 + 8H_2O$$

$$2MnO_4^- + 10Fe^{2+} + 8H^+ \rightleftharpoons 10Fe^{3+} + 2MnSO_4\ (acidic\ medium)$$

$$Mn^{7+} + 5e \rightarrow Mn^{2+}\ \ (5\,e\ gain - reduction)$$

$$Fe^{2+} \rightarrow Fe^{3+} + e\ \ (1\,e\ loss - oxidation)$$

$$Eq.g\,(\,KMnO_4\,) = \frac{MM}{5} = \frac{157.9}{5} = 31.6$$

The relationship between the various concentration units.

The relationship between molar concentration and mass concentration:

$$C_M(\text{ molar concentration}) = \frac{n}{V} \quad \text{and} \quad C_m(\text{mass concentration }) = \frac{m(\text{ solute})}{V(\text{solution})}$$

$$n = \frac{m}{M}$$

$$\Rightarrow \quad C_M = \frac{1}{M}\frac{m}{V} \quad \Rightarrow$$

The relationship between molar concentration and regularity:

$$N = \frac{n\ eq.g}{V}\ \frac{\text{عدد المكافئات الغرامية}}{\text{الحجم}} \qquad \textit{and}$$

$$n\ eq\ g = \frac{r}{eq.}$$

$$\Rightarrow \qquad N = \frac{m}{eq\ g\ .V}$$

$$N = \frac{m}{eq\ g.V} \qquad \text{and} \qquad eq\ g = \frac{M}{z} \qquad \Rightarrow \qquad \boxed{N = \frac{m.z}{M.V}}$$

$$N = \frac{m.z}{M.V} \quad = \quad n.\frac{z}{V} \quad = \quad C.z \qquad \Rightarrow \qquad \boxed{N = C_M.z}$$

Exercice: We dissolve a mass m = 20 g of phosphoric acid H_3PO_4 in a volume of 250 cm^3, and solution S_1 is formed.

1. Calculate the number of moles of acid dissolved
2. Calculate the molar concentration of solution S1
3. Calculate the mass concentration of solution S1
4. Calculate the gram equivalent of phosphoric acid and then deduce the regularity of S1
5. What is the volume of distilled water that must be added to solution S1 to obtain a normality of N_2 =0.5

Molality : is calculated by the following mathematical relationship:

$$\boldsymbol{Molality\ (m) = \frac{number\ of\ moles\ of\ solute\ (mol)}{mass\ of\ solvent\ (\ Kg)}}$$