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COMMON BASE DEPARTMENT

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Elements of chemistry

Chapter 1

Fundamental concepts

I. Definition of matter

II. Changes of state of matter

III. Classification of matter

IV. Notion of atom, molecule, mole and Avogadro number

V. Law of conservation of mass (Lavoisier), chemical reaction

VI. Qualitative and quantitative aspects of matter

I. Definition of Matter

Matter is everything that has mass and occupies volume in space.Matter can exist in three different physical states:

- **The solid state**: has a defined volume and shape.
- **The liquid state**: with a defined volume but no precise shape, it takes on the form ofits container.
- **The gaseous state**: have no defined volume or shape, taking on the volume and shapeof its container.

II. Changes in the state of matter

Changes of state are significant physical changes that occur at temperatures that are characteristic of the substance.

Example: Melting temperature of water: 0 °C.

Copper melting temperature: 1084 °C.

II. a. Physical change

A physical change is a transformation that does not change the nature of a substance; it simply involves a change in its physical state, shape or dimensions.

Depending on external conditions (temperature and pressure), the same substance can be solid, liquid or gaseous. In fact, the transition of matter from the solid to the liquid state takes place by melting, from the liquid to the gaseous state by vaporization, and from the solid to the gaseous state by sublimation.

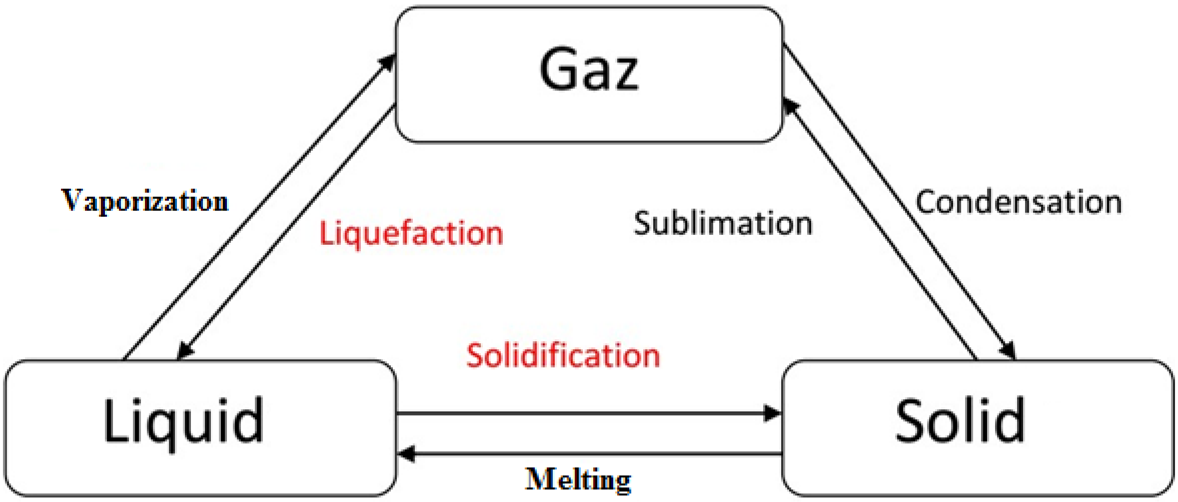


Figure I.1: Changes of state

II.b. Chemical change

A chemical change is a transformation that changes the nature of a substance by means of a chemical reaction.

Example: Corrosion: iron gives rust.

Combustion: wood burns to produce ash and gases. Chemical changes can be recognized by certain indicators:

- Gas formation
- Precipitate formation
- Color change
- Energy production in the form of light and heat.

III. Matter classification

Matter is found in the form of **pure bodies** (simple or compound) and **mixtures** (homogeneous or heterogeneous).

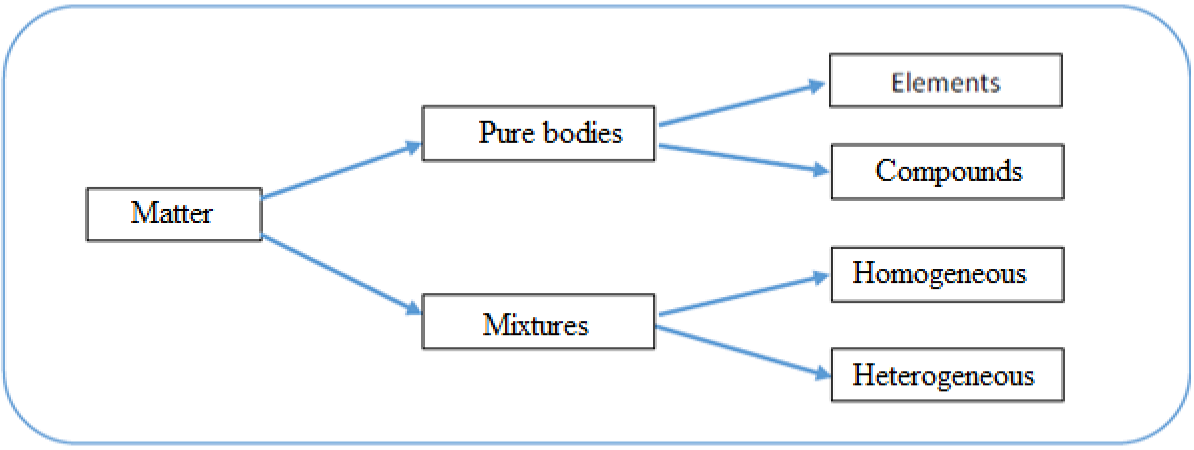


Figure I.2: Matter

- ✓ A pure body is one made up of a single chemical entity (atom, ion or molecule).
- Simple pure bodies: made up of molecules composed of a single type of atom (H_2 ,O_2 ,Fe, etc...).
- Compound pure bodies: made up of molecules with different atoms (H_2O, NaCl, CO_2, etc.).

- ✓ Mixtures are compounds whose molecules are different (simple bodies or compounds). They fall into two categories:
- Homogeneous mixtures: have the same appearance (same properties) at all points and within which it is impossible to distinguish several constituents. They form a single phase.

Examples: salt or sugar water (solutions), air, steel (alloys), etc...

- Heterogeneous mixtures: parts with different appearances (two or more phases) can be visually or microscopically distinguished.

Example: water-oil mixture, sand, unfiltered natural water, etc...

IV. Notion of atom, molecule, mole and Avogadro number

IV-1. Atoms:

Matter is made up of elementary particles called atoms. An atom is considered the smallest particle of an element and is, according to Greek etymology, indivisible and indestructible. There are 111 species of atom, each differing from the next in structure, mass and physico-

chemical properties. Each element is designated by an abbreviation or symbol noted : $^{A}_{Z}X$

where : **A**: mass number; **Z**: charge number (atomic).

IV-2. Molecules

A molecule is a union of two or more atoms linked together by bonds. It is the smallest part of a compound that has the same properties as the compound itself.

Examples: H_2O, H_2, HCl, H_2SO_4, etc....

IV-3. Mole (unit of amount of substance) and Avogadro number (N_A):

The mole is the quantity of matter of a system containing **fi** identical entities. In practice, this number N (also noted: $\mathbf{N_A}$) is called **Avogadro**'s number and is worth approximately $\mathbf{6.023.10^{23}}$ $\mathbf{mol^{-1}}$. The mole is also defined as the number of carbon 12 atoms contained in 12 g of carbon 12 (^{12}C). It applies to all elementary species: atoms, particles (electrons, protons, etc), molecules, ions.

Example: One mole of atom corresponds to $6.023.10^{23}$ atoms

One mole of electrons corresponds to $6.023.10^{23}$ electrons

Example:

Determine the number of atoms contained in 12 g of carbon, knowing that the mass of a carbon 12 atom is $1.9926.10^{-26}$ Kg**.** (This mass was determined using a mass spectrometer).

Solution:

1 mol (of carbon 12) $\longrightarrow$ 12g $\longrightarrow$ N atoms

$1.9926 \,.\, 10^{-26}$ Kg $\longrightarrow$ 1 atom

$$1\text{mol} = \frac{12}{1{,}9926.10^{-26}.10^{3}} = 6.023.10^{23}$$

IV-4. Atomic molar mass:

The atomic mass of an element is the mass of one mole of atoms, also called the atomic molar mass, or the mass of **NA** atoms.

Example: 1 mol Na atoms → **NA** Na atoms → 23g

The mass of a Na atom is: 1 Na atom → m_{Na}

Therefore: $m_{Na} = 23/ \mathbf{N_A} = 3.8.10^{-23}$ g

IV-5. Molecular mass:

This is the mass of one mole of a molecule. It is equal to the sum of the molar masses of the atoms making up the molecule.

Example:

The $Zn(OH)_2$ molecule . The mass of one mole is: $M(Zn(OH)2\ ) = M(Zn) + 2.M(O) + 2.M(H)$ = 65.37+2.16+2.1 = 99.37 g/mol.

The mass of a $Zn(OH)_2$ molecule : 1 $mol\ de$ Zn(OH)2 $\rightarrow$ 99.37 $g \rightarrow 6.023\ .10^{23}$

molecules mZn(OH)2 $\rightarrow$ 1 molecul

$mZn(OH)2 = (99.37/6.023\ .10^{23}\ ) = 1.649.10^{-22}$ g

IV-6. Molar volume:

This is the volume occupied by one mole of substance. In the case of gases, and under normal conditions of pressure and temperature (n = 1mol; T = 0°C; P = 1 atm): **V = 22.4 l**

Example:

$5.38.10^{19}$ α particles (He^{2+}) lead to 2 cm^3 of helium gas under standard temperature and pressure conditions. Determine the Avogadro number.

Solution: 1 mol of (He) $\rightarrow$ 22.4 l $\rightarrow$ N atoms

2.10^{-3} l $\rightarrow 5.38\ .10^{19}$ atoms

$N = (22.4.5{,}38.10^{19} / 2.10^{-3}\ ) = 6.025\ .10^{23}\ atomes$

IV- 7. Unit of atomic mass (u.m.a):

The atomic mass unit is defined as the fraction (1/12) of the mass of a carbon 12 atom (^{12}C).

1 u.m.a = (1/12).(mass of an atom of ^{12} C)

1 u.m.a = (1/12).(12/6.023.10^{23}) = 1.666.10^{-24} g

1 u.m.a = 1.666.10^{-27} kg

V. Law of conservation of mass (Lavoisier), chemical reaction

V- 1. In a laboratory (macroscopic level):

In a chemical reaction taking place in a closed environment, the sum of the masses of the reactants is always equal to the sum of the masses of the products formed.

Example: $A + B \longrightarrow C + D$

mass of A + mass of B = mass of C + mass of D

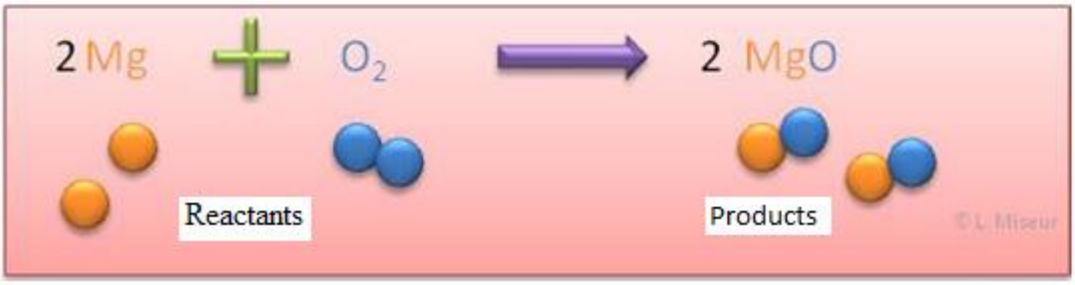
"Nothing is lost, nothing is created, everything is transformed".

Note: The law of conservation of mass could not be used to explain open systems or certain reactions.

V- 2. At the level of molecules and atoms (atomic level):

Lavoisier's law implies, at the atomic level, that before the chemical reaction and after the chemical reaction, there must be the same number of atoms in the system.

The chemical reaction is the reorganization of the atoms of the reactants into new molecules = the products (which have a different nature and therefore different physical and chemicalproperties.



VI. Qualitative and quantitative aspects of the matter

VI- 1. Solutions

A solution is a homogeneous mixture of two or more components. In liquid, gaseous or solid phase.

A solution is obtained by dissolving a chemical species in a liquid called a solvent. The dissolved species, called the solute, may be ionic (cation, anion) or molecular. **solution = solvent + solute**

- A solvent is any liquid substance that has the power to dissolve other substances.
- A solute is a chemical species (molecular or ionic) dissolved in a solvent. The solvent is always in much greater quantity than the solute(s).
- This homogeneous mixture (solvent + solute) is called an aqueous solution if the solvent is water.

VI- 2. Dilution

Diluting an aqueous solution involves reducing its concentration by adding solvent. The initial solution with a higher concentration is called the **stock solution**, and the solution after dilution is called the **diluted solution**. After dilution, the quantity of substance is the same.

We can write : $C_m = \frac{n}{V} \Rightarrow n = C_m * V$

Before dilution : $n_1 = C_1.V_1$

After dilution: $n_2 = C_2.V_2$; Number of moles is the same:

$$n_1 = n_2 \Rightarrow C_1.V_1 = C_2.V_2$$

VI- 3. Saturation

A solution is said to be saturated if the maximum quantity of solute to be dissolved by the solvent is exceeded. A precipitate appears at the end of dissolution.

VI- 4. The mole number

The mole number is the ratio of the compound's mass to its molar mass.

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

n: number of moles, m: mass of compound; M: molar mass of compound

VI- 5. Molar concentration or molarity

The molar concentration of a chemical species in solution is the quantity of that species per liter of solution. For a species i, it is denoted C_i or [i] and its unit is the mole per liter ($mol.L^{-1}$).

$$C_i = \frac{n_i}{V} = [i]$$

With: n_i number of moles (mol), V: volume in liters (L), C_i or concentration in moles per liter($mol.L^{-1}$ **).**

VI- 6. Concentration by weight or mass

Mass concentration is the ratio between the mass of the compound and the volume of the solution.

$$C_m = \frac{m_i}{V} \qquad (g.L^{-1}).$$

With: m_i mass of solute i (g), V: volume in liters (L), C_m mass concentration.

However, these two expressions of concentration are linked: knowing one, it's possible to deduce the other.

In fact, $m_i = M_i \times n_i \Leftrightarrow C_m = \frac{m_i}{V} = \frac{M_i \times n_i}{V}$. We then find the expression for the molar concentration *C* of solute "i". Thus, $C_m = C_i \cdot M_i$

VI- 7. Molality

The molality of a component i is defined as the quantity of i relative to the mass of solvent. Its unit is the mole per kilogram ($mol.Kg^{-1}$) and it is noted b_i.

$$b_i = \frac{n_i}{m_{solvant}}$$

Where n_i : quantity of solute *i* and $m_{solvent}$: mass of solvent (Kg).

VI- 8. Normal concentration or normality

The **normality** of a solution is the number of gram equivalents of solute contained in one liter of solution. This measure of concentration is inseparable from a particular chemical reaction (acid-base or oxidation-reduction reaction), which is implicitly considered.

The unit of normality is the gram equivalent per liter, represented by the symbol **N** :

1N = 1 (eq.g/l). Given the definition of equivalent mass, the normality of a solution will always be an integer multiple (1, 2, 3, ...) of its molarity, i.e. : $\mathbf{N = z.C_i}$

Gram equivalent (Eq.g): the ratio of the mass of the solute in its pure state to its equivalent molar or atomic mass.

Equivalent molar or atomic mass ($\mathbf{M_{eq\text{-}gram}}$): this is the molar or atomic mass of the element considered in relation to the number of electrons, proton or hydroxyl exchanged during a reaction.

$$N = \frac{n_{eq-gramme\ de\ soluté}}{V_{solvant}}$$ With : $$n_{eq-gramme\ de\ soluté} = \frac{m}{M_{eq-gramme}}$$

And $$M_{eq-gramme} = \frac{M}{Z}$$

Meq-gram: equivalent mass and M is the molecular mass of the solute.
Z represents the number of H^+ protons in the case of an acid solute and the number of hydroxides OH^- in the case of a basic solute.

VI- 9. Molar and mass fractions

When a phase (liquid, solid or gas) contains several chemical species, it may be practical to express the proportion of each entity within the phase. In this case, we use the concepts of mass or molar fractions.

- ✓ The **mole fraction** x of a compound (A) in a mixture containing (A) and (B) is expressed in terms of its quantity of subst $x_A = \frac{n_A}{n_A + n_B}$

Generalizing to a mixture containing "i" compounds : $$x_i = \frac{n_i}{\sum_i n_i}$$

Note that for gaseous mixtures, the mole fraction is given by *y*.

- ✓ The **mass fraction** of compound (A) in the mixture (A) and (B) is expressed by the following relationship : $\bar{x}_A = \frac{m_A}{m_A + m_B}$

This formula can be generalized for a mixture containing "i" compounds: $$\bar{x}_i = \frac{m_i}{\sum_i m_i}$$

VI- 10. Title

The title of a solution is the mass, in g, of the solution in 1 ml of solvent. It is given by the following relationsh $T_{A/B} = \frac{m_A}{V_B}$

VI- 11. Molar volume

This is the volume occupied by one mole of substance in the gaseous state. Under standard conditions of pressure and temperature (SCPT): P = 1atm, T= 0 °C = 273K, one mole of gaseous substance occupies a volume of 22.4 L, noted V_m .

For a number of moles n, the volume is: $\mathbf{V = V_m \cdot n}$

Where : V is the volume of the substance, n is the amount of substance (mole) and V_m is themolar volume ($L.mol^{-1}$).

VI- 12. Density and specific gravity (volumic mass)

- **Volumic mass**

The specific gravity (volumic mass) is a physical quantity that characterizes the mass of a material per unit volume. It is determined by the ratio: $\rho = \frac{m}{V}$ (Kg/m^3), (g/cm)^3

where m is the mass of the homogeneous substance occupying a volume V.

- **Density**

- **Solid-liquid density**

The density of a solid or liquid body is the ratio of the mass of a certain volume of this body to the same mass of volume of water. It is given by the following relationship:

$$d = \frac{\rho_{corps\ (S\ ou\ l)}}{\rho_{eau}} = \frac{m/V}{m'/V} = \frac{m}{m'}$$

- **Gas density**

In the case of gas or steam, the gaseous reference body is air, at the same temperature and pressure.

$$d = \frac{masse\ d'un volume\ V\ du\ gaz}{masse\ du\ même\ volume\ V\ d'air} = \frac{M_{gaz}}{M_{air}} = \frac{M_{gaz}}{29} \quad ; \quad M_{gaz} = d_{gaz} * 29$$