CHAPTER 1

FUNDAMENTAL NOTIONS

Chapter contents

- **Macroscopic Characterization and States of the Matter.**
- **Changing States of Matter**
- **Atoms, Elements and Compounds**
- **Mole and Avogadro number**
- **Atomic Mass Unit**
- **Atomic and Molecular Molar Mass, Molar Volume,**
- **Weight Law: Conservation of Mass (Lavoisier), Chemical Reaction, Qualitative and Quantitative Aspect of Matter.**

CHAPTER 1. FUNDAMENTAL NOTIONS

I- Macroscopic characterization and states of the matter

The physical properties of matter can be viewed from either the macroscopic and microscopic level. The macroscopic level includes anything seen with the naked eye and the microscopic level includes atoms and molecules, things not seen with the naked eye. Both levels describe matter. Matter is anything that occupies space and has mass.

- On the **microscopic level**, chemical symbols represent the behavior of individual atoms and molecules. Atoms and molecules are much too small to be seen, but we can nevertheless describe their microscopic behavior if we read the equation between them.

For example “Two molecules of hydrogen react with one molecule of oxygen to yield two molecules of water.” It’s often helpful to visualize a molecule as a collection of spheres stuck together.

- On the **macroscopic level**, formulas and equations represent the large-scale behaviors of atoms and molecules that give rise to visible properties. In other words, the symbols and represent not just single molecules but vast numbers of molecules that together have a set of measurable physical properties.

A single isolated molecule is neither solid nor liquid nor gas, but a huge collection of molecules appears to us as a colorless liquid that freezes at 0°C and boils at 100°C. Clearly, it’s this macroscopic behavior we deal with in the laboratory when we weigh out specific amounts of reactants, place them in a flask, and observe visible changes.

Moreover, the units of measurement for these properties are also different; for macroscopic properties, the unit of measurement is on a scale that is visible to the naked eye. This includes centi-, kilo-, mega-, etc. For microscopic properties, the unit of measurement is on a scale that is invisible to the naked eye and includes milli-, micro-, nano-, pico-, etc.

Under an international agreement concluded in 1960, scientists throughout the world now use the International System of Units for measurement, abbreviated IS in the macroscopic level of units. Based on the metric system, which is used in all industrialized countries of the world except the United States, the SI system has seven fundamental units (Table 1.1).

Table 1.1. The Seven Fundamental SI Units of Measure

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Temperature	kelvin	K
Amount of substance	mole	mol
Time	second	s
Electric current	ampere	A
Luminous intensity	candela	cd

Table 1.2. The measurement units for macroscopic and microscopic properties.

Physical Quantity	Unit name	Unit Symbol	Definition
Length	Angstrom	Å	10^{-10}m
	nanometer	nm	10^{-9} m
Area	square meter	m^2	
Volume	cubic meter	m^3	dm^3, 10^{-3} m^3
	liter	ℓ	
	cubic centimeter	cm^3, mℓ	
Mass	atomic mass unit	U(a.m.u)	1.66054×10^{-27} kg
	microgram	µg	10^{-9} kg (10^{-6} g)
Density	Kilogram per cubic meter (IS)	Kg/m^3	
	Gram per milliliter or gram per cubic centimeter	g/mℓ or g/cm^3	
Force	newton	N	kg . m/s^2
Pressure	pascal (SI)	Pa	N/m^2
	bar	bar	10^5 Pa
	atmosphere	atm	101 325 Pa
	torr (millimeter of mercury)	torr (mm Hg)	atm/760 or 133.32 Pa

Exercise 1:

A benzene solution containing 0.10 mm^3 of stearic acid is dropped into a tray full of water. The acid is insoluble in water but spreads on the surface to form a continuous film of area 400 cm^2

after all of the benzene has evaporated. What is the average film thickness in (a) nanometers, (b) angstroms?

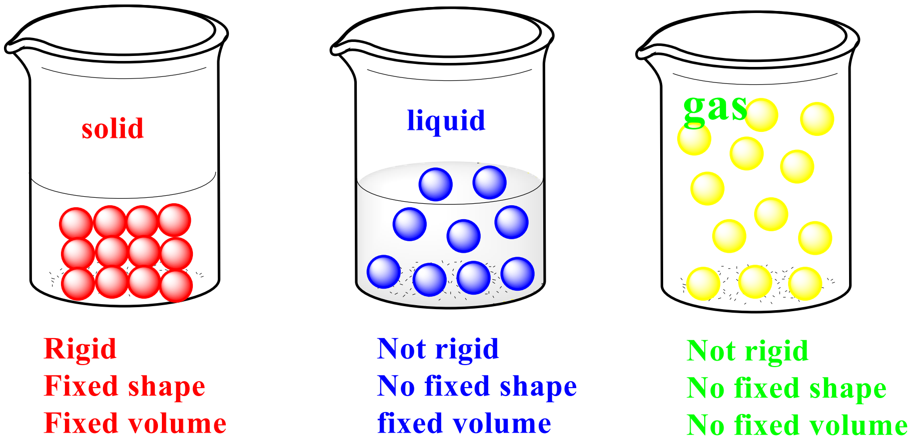
II- Changing states of matter

The state of matter of a substance depends on how fast its particles move and how strong the attraction is between its atoms and molecules. All matter can exist in three states: solid, liquid and gas.

Solids: keep their shape and volume. The particles of the substance vibrate in place. The vibration isn't strong enough to overcome the attraction of the particles and cause them to separate. As a result, the forces between the particles cause them to lock together.

Liquids don't have a shape of their own. They take on the shape of the container they are in. Liquids do have a definite volume, though. The particles of a liquid move faster than particles of a solid. As a result, the particles in a liquid can overcome some of the attraction between them. Unlike the particles in a solid, which are locked together, the particles in a liquid can flow around and over each other.

Gases can flow throughout a room. Particles in a gas move so fast they are able to overcome the attraction between them. The particles of a gas will drift apart and will spread out in all directions. They do this whether they are filling up a balloon, a room, or all the Earth's atmosphere.



Changing the state of matter of a substance is a physical change. It is usually caused by changing the temperature or surrounding pressure of a substance.

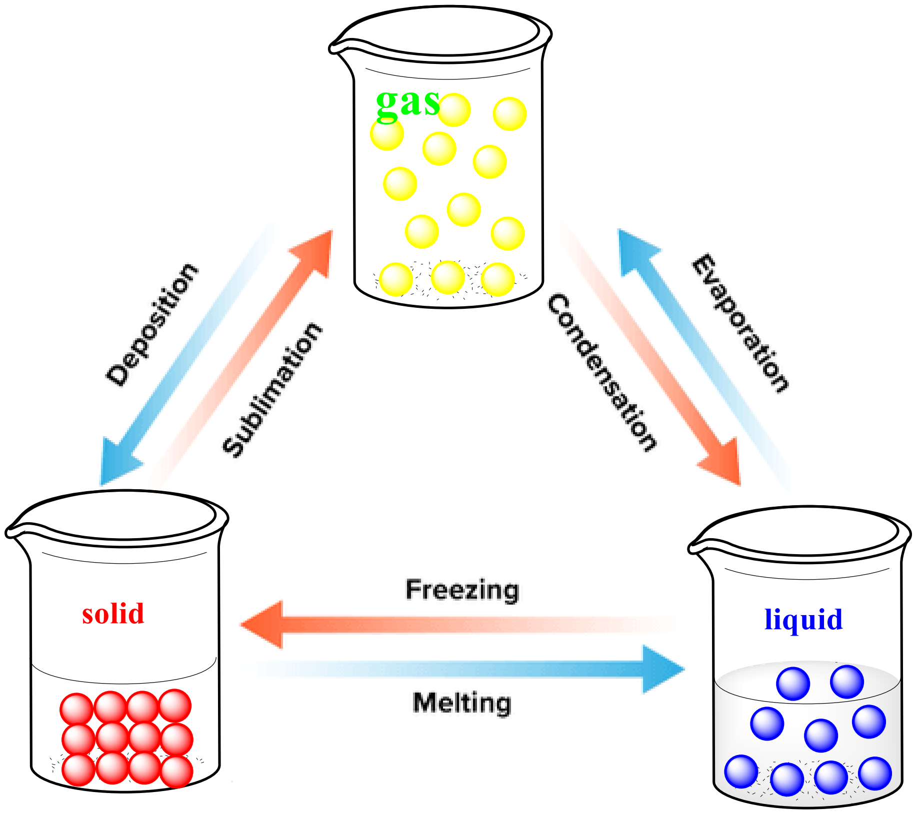
Solid, liquid and gas **expand when heated**. They **contract when cooled**. The increase in volume of matter due to an increase in temperature is called **thermal expansion**.

When heat is added to a solid, its temperature will rise to a certain point where the solid starts to melt. This point is called the **melting point**. When heat is removed from the liquid, its temperature drops to a certain point where the liquid starts to freeze. This point is called the **freezing point**.

The temperature at which a liquid changes into a gas is called the **boiling point**.

The change of state from a liquid to a gas is called **evaporation**. The change of state from a gas to a liquid is called **condensation**.

If solids have enough vapor pressure at a particular temperature then they can change directly into air. The direct change of state from solid to gas is called **sublimation**.



III- Atoms, elements and compounds

Atoms: The basic building block of all matter is called an atom.

Atoms are a collection of various subatomic particles containing negatively charged electrons, positively charged protons and neutral particles called neutrons. Each atom has its own unique number of protons, neutrons and electrons.

Examples: ${}^{16}_{8}O$, ${}^{26}_{56}Fe$, ${}^{11}_{23}Na$.

The Elements: is a substance that cannot be decomposed into simpler substances by chemical changes.

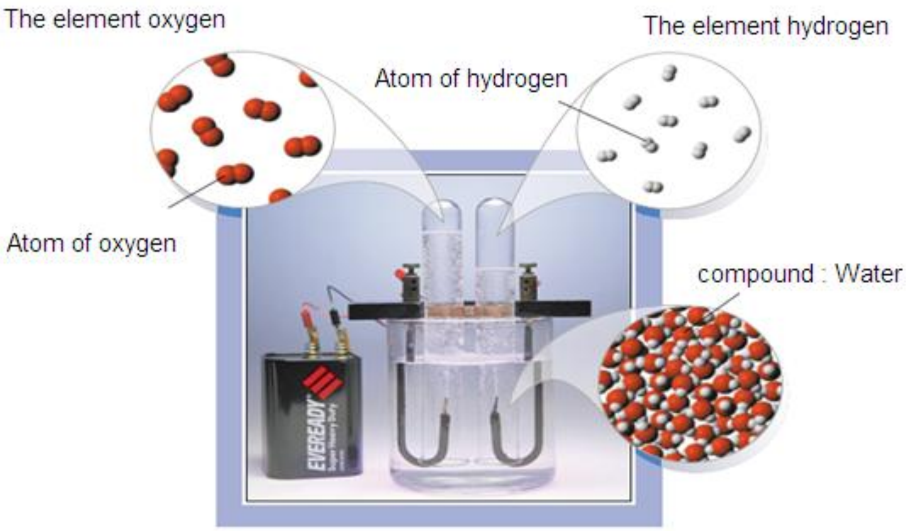
- At present, **118 elements** have been positively identified.

Table 1.3. Some Common Elements and Their Symbols

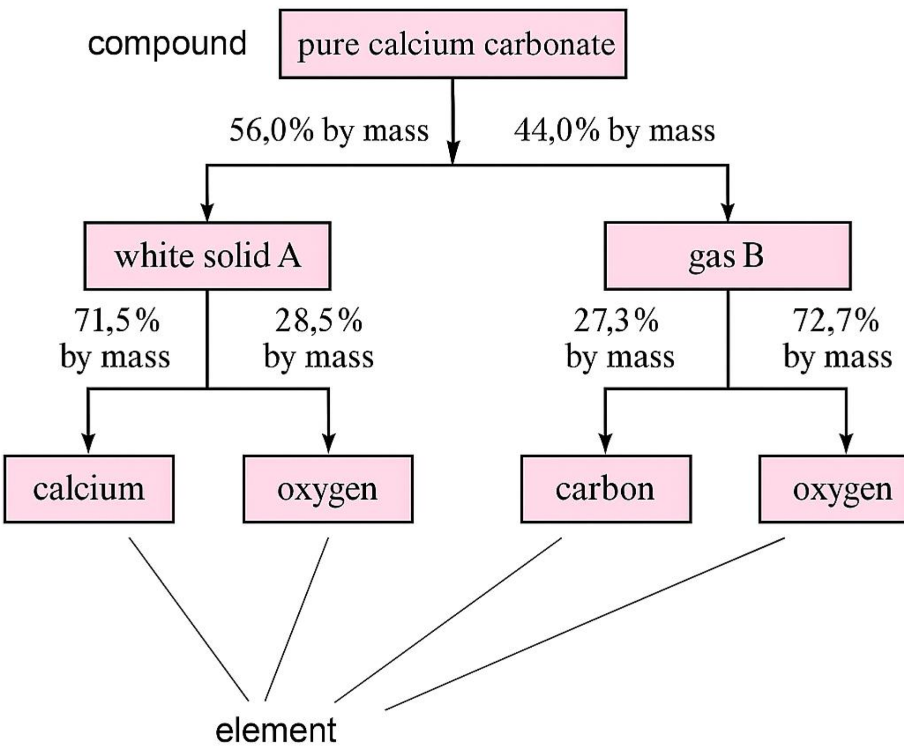
Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Phosphorus	P	Barium	Ba
Hydrogen	H	Platinum	Pt	Bromine	Br
Iodine	I	Potassium	K	Calcium	Ca
Iron	Fe	Silicon	Si	Carbon	C
Lead	Pb	Silver	Ag	Chlorine	Cl
Magnesium	Mg	Sodium	Na	Chromium	Cr
Mercury	Hg	Sulfur	S	Cobalt	Co
Nickel	Ni	Tin	Sn	Copper	Cu
Nitrogen	N	Zinc	Zn	Gold	Au

Samples of mercury, silver, and sulfur

A compound: is a substance that can be decomposed by chemical means into simpler substances, always in the same ratio by mass.



Example:



- All the many kinds of matter can be classified as either pure substances or mixtures. Pure substances, in turn, can be either elements or chemical compounds.

Chemical Compounds: A chemical compound is a pure substance that is formed when atoms of two or more different elements combine and create a new material with properties completely unlike those of its constituent elements.

For example, when atoms of sodium (a soft, silvery metal) combine with atoms of chlorine (a toxic, yellow-green gas), the familiar white solid called sodium chloride NaCl (table salt) is formed.

Mixtures: A mixture is a combination of two or more substances in which the substances retain their distinct identities.

Some examples: air, soft drinks, milk, and cement.

- Mixtures can be further classified as either **heterogeneous** or **homogeneous**.

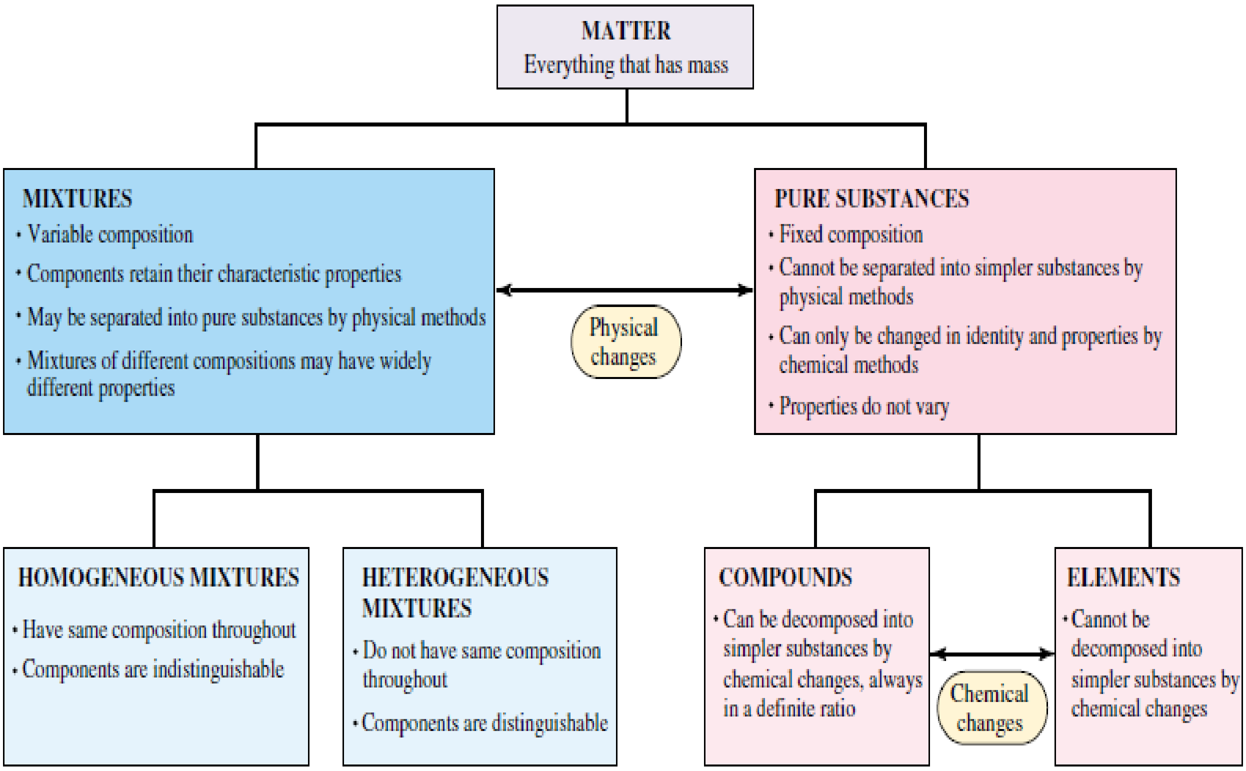
Heterogeneous mixtures are those in which the mixing is not uniform and which therefore have regions of different composition.

Examples: Sand with sugar, water with gasoline, and dust with air.

Homogeneous mixtures are those in which the mixing is uniform and which therefore have a constant composition throughout.

Examples: Air is a gaseous mixture of (primarily) oxygen and nitrogen, seawater is a liquid mixture of (primarily) sodium chloride dissolved in water, and brass is a solid mixture of copper and zinc.

- With liquids it's often possible to distinguish between a homogeneous mixture and a heterogeneous one simply by looking. Heterogeneous mixtures tend to be cloudy and will separate on standing, whereas homogeneous mixtures are often transparent.

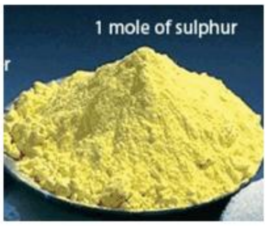


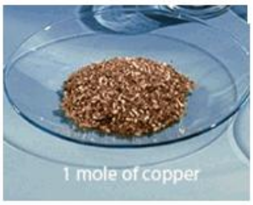
IV- Mole and Avogadro number

The concept of a mole as applied to atoms is especially useful. It provides a convenient basis for comparing the masses of equal numbers of atoms of different elements.

Table 1.4. Mass of one mole of atoms of some common elements.

Element	A sample with a mass of	Contains
Carbon	12 g of C	6.02×10^{23} C atoms or 1 mol of C atoms
sulfur	32.1 g S_8	6.02×10^{23} S atoms or 1 mol of S atoms
Gold	197 g of Au	6.02×10^{23} Au atoms or 1 mol of Au atoms
Hydrogen	1 g of H_2	6.02×10^{23} H atoms or 1 mol of H atoms





The number of atoms in a 12 g sample of carbon 12 is called Avogadro's number (to which we give the symbol N). Recent measurements of this number give the value **6.0221367×10^{23}**.

Exercise 2:

How many molecules are there in 20.0 g of benzene, C_6H_6?

V- Atomic mass unit

The mass of a single atom in grams is much too small a number for convenience, and chemists therefore use a unit called an atomic mass unit (amu), also known as a dalton (Da). One amu is defined as exactly one twelfth the mass of an atom of $^{12}_{6}C$ and is equal to the value of :

$$\mathbf{1.660\,539 \times 10^{-24}\ g = 1.660539 \times 10^{-27}\ kg.}$$

$$\begin{cases} \text{One mole of carbon} \rightarrow\ N \text{ atoms of carbon} \rightarrow\ 12\text{ g} \\ \text{One atom of carbon} \rightarrow\ \text{mass of one atom of carbon} \end{cases} \Rightarrow \text{mass of one atom of carbon} = \frac{12}{N}$$

$$1\text{ amu} = \frac{1}{12} \times \frac{12}{N} = \frac{1}{N} \Rightarrow \mathbf{1amu = 1.66 \times 10^{-24}\ g}$$

Exercise 3.

- How many moles of Zinc are there in a sample of 23.3 g of Zn?
- How many atoms of Zn are there in this sample?
- What is the mass (in grams and in amu) of 4.52×10^{23} Zn atom?

VI- Atomic and molecular molar mass, molar volume

1- Atomic molar mass

An element's atomic mass is the weighted average of the isotopic masses of the element's naturally occurring isotopes.

Example: Carbon occurs on Earth as a mixture of two isotopes, ${}_6^{12}C$ (12 amu, 98.89% natural abundance) and ${}_6^{13}C$ (13.0034 amu, 1.11% natural abundance) (A third carbon isotope, ${}_6^{14}C$, its an abundance is so low that it can be ignored when calculating atomic mass). The average isotopic mass—that is, the atomic mass—of a large collection of carbon atoms is 12.011 amu.

2- Molecular molar mass

Ethanol, whose molecular formula is C_2H_5OH, has a molecular mass of 46.1 amu and a molar mass of 46.1 g/mol.

The molecular molar mass of a substance is the mass of one mole of the substance. Carbon-12 has a molar mass of exactly 12 g/mol, by definition. For all substances, the molar mass in grams per mole is numerically equal to the formula mass in atomic mass units.

Exercise 4:

What is the mass in grams of a chlorine atom, Cl? b. What is the mass in grams of a hydrogen chloride molecule, HCl?

3- molar volume

In 1811, Amedeo Avogadro postulated that at the same temperature and pressure, equal volumes of all gases contain the same number of molecules. This Law can also be stated as follows: At constant temperature and pressure, the volume, V, occupied by a gas sample is directly proportional to the number of moles, n, of gas.

$$V = k\,n \quad \text{or} \quad \frac{V}{n} = k \quad \text{(constant P, T)}$$

For two samples of gas at the same temperature and pressure, the relation between volumes and numbers of moles can be represented as:

$$\frac{V_1}{n_2} = \frac{V_2}{n_2} \quad \text{(constant P, T)}$$

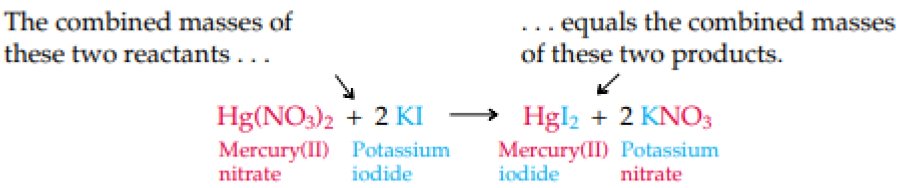
The standard molar volume of an ideal gas is taken to be **22.414 liters per mole** at constant P, T.

I- Weight law:, Conservation of mass, chemical reaction, Qualitative aspect of matter, Quantitative aspect of matter

1- Conservation of mass Law (Lavoisier)

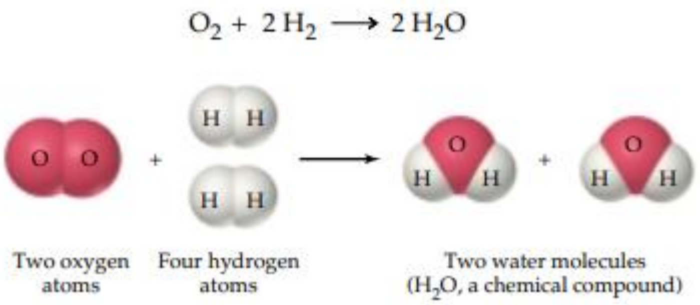
Lavoisier demonstrated by careful measurements that when combustion is carried out in a closed container, the mass of the combustion products exactly equals the mass of the starting reactants.

LAW OF MASS CONSERVATION Mass is neither created nor destroyed in chemical reactions. That means: the total mass remains constant during a chemical change.



2- A chemical reaction

A chemical equation is written in a standard format, in which the starting substances, or **reactants**, are on the left, the final substances, or **products**, are on the right, and **an arrow** is placed between them to indicate **a transformation**. The numbers and kinds of atoms are the same on both sides of the reaction arrow, as required by the law of mass conservation.



A chemical equation also indicates the relative amounts of each reactant and product in a given chemical reaction. Then the equation can be written :

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \rightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{mol} & & 1\text{mol} & & 2\text{mol} \end{array}$$

This interpretation tells us that one mole of methane reacts with two moles of oxygen to produce one mole of carbon dioxide and two moles of water.

A balanced chemical equation may be interpreted in terms of moles of reactants and products and Never start a calculation involving a chemical reaction without first checking that the equation is balanced.

Exercise 5:

1- How many moles of water could be produced by the reaction of 3.5 mol of methane with excess oxygen (i.e., more than a sufficient amount of oxygen is present).

2- Balance the following equations :

(a) $KNO_3 \rightarrow KNO_2 + O_2$

(b) $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$

(c) $2Na + 2H_2O \rightarrow 2NaOH + H_2$

(d) $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

(e) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

(f) $Fe_3O_4 + 4H_2 \rightarrow 3Fe + 4H_2O$

(g) $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

(h) $Zn + 2KOH \rightarrow K_2ZnO_2 + H_2$

(j) $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

(k) $Al(NO_3)_3 + 3NH_3 + 3H_2O \rightarrow Al(OH)_3 + 3NH_4NO_3$

(l) $Al(NO_3)_3 + 4NaOH \rightarrow NaAlO_2 + 3NaNO_3 + 2H_2O$

Chemical equations describe reaction ratios, that is, the mole ratios of reactions and products as well as the relative masses of reactants and products.

We use the coefficients to get the mol ratio of any two substances we want to relate then we apply it as:

Moles of desired substance = (moles of substance given) × (mole ratio from balanced chemical equation)

Mole substance must appear in the numerator of the mole ratio, i.e., do not just write $\frac{\text{mol}}{\text{mol}}$, write $\frac{\text{mol of something}}{\text{mol of something else}}$ and giving the formulas of the two substances involved.

Exercise 6:

What mass of oxygen is required to react completely with 1.2 mol of CH_4?

- **The Limiting Reactant Concept**

The substance that was used up first, called the **limiting reactant**, the other reactant is said to be the **excess reactant**. The limiting reactant is not necessarily the reactant present in the smallest amount.

Exercise 7:

What mass of CO_2 could be formed by the reaction of 16 g of CH_4 with 48 g of O_2?

- **Percent Yields from Chemical Reactions**

The theoretical yield from a chemical reaction is the yield calculated by assuming that the reaction goes to completion.

$$\textbf{Percent yield (\%)} = \frac{\textbf{actual yield of product}}{\textbf{theoretical yield of product}} \times 100$$

Exercise 8:

A 15.6 g sample of C_6H_6 is mixed with excess HNO_3. We isolate 18 g of $C_6H_5NO_2$. What is the percent yield of $C_6H_5NO_2$ in this reaction

3- Qualitative aspect of matter

Qualitative Property: is a characteristic of a substance that can be described but not measured. You can describe matter using the following terms.

- **To describe all substances**:
  - ✓ clarity : transparent (clear), translucent (cloudy), opaque
  - ✓ colour : colourless, red, orange, blue, white, etc
  - ✓ odour : odourless, sweet, sour, burnt, aromatic, fragrant, nauseating, sharp, acrid, choking
- **To describe solids only**:
  - ✓ texture : crystalline, granular, waxy, flaky
  - ✓ lustre : shiny, dull, metallic, greasy, glassy
  - ✓ hardness : hard, soft, flexible, brittle
- **To describe liquids** only: viscosity

4- Quantitative aspect of matter

Quantitative Property: is a characteristic of a substance that can be measured numerically.

Stoichiometry: Description of the quantitative relationships among elements and compounds as they undergo chemical changes. It is the calculation of the quantities of reactants and products involved in a chemical reaction

Reaction stoichiometry: Description of the quantitative relationships among substances as they participate in chemical reactions.

- ❖ Such calculations are fundamental to most quantitative work in chemistry

Density : is the mass of an object divided by its volume:

$$\textbf{density} = \frac{\textbf{mass}}{\textbf{volume}} \Rightarrow \textbf{d} = \frac{\textbf{m}}{\textbf{V}}$$

Where d, m, and V denote density, mass, and volume, respectively

The SI-derived unit for density is the **kilogram per cubic meter (kg/m^3)**. Therefore, **grams per cubic centimeter (g/cm^3)** and its equivalent, **grams per milliliter (g/mL)**, are more commonly used for solid and liquid densities.

Percentage Composition: A chemical formula may be used to compute the percentage composition of a compound; that is, the percentage by weight of each type of atom in the compound.

Solution concentration is sometimes expressed in terms of the mass percentage of solute—that is, the percentage by mass of solute contained in a solution.

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

Molarity: A solution that contains one mole of solute per liter of solution is known as a one molar solution; it is abbreviated 1.00 M. In general

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

Molality: The molality of a solution is the moles of solute per kilogram of solvent.

$$\text{molality of solution} = N = \frac{\text{moles of solute}}{\text{kilograms of solution}}$$

Problem:

- A piece of gold ingot with a mass of 0.301 kg has a volume of 15.6 cm^3. Calculate the density of gold.
- Calculate the percentages of oxygen and hydrogen in water, H_2O.
- How would you prepare 425 g of an aqueous solution containing 2.40% by mass of sodium acetate, $NaC_2H_3O_2$?
- A compound contains 90.6% Pb and 9.4% O by weight. Find the empirical formula.
- What volume of 0.250 M Na_2CrO_4 will be needed in order to obtain 8.10 g of Na_2CrO_4?
- What is the molality of a solution containing 5.67 g of glucose dissolved in 25.2 g of water?
- A solution is prepared by dissolving 0.131 g of a substance in 25.4 g of water. The molality of the solution is 0.056 m. What is the molecular mass of the substance?

References

1- Darrell D. Ebbing, Steven D. Gammon; General Chemistry, Ninth Edition., **2009** by Houghton Mifflin Company.

2- McMurry Fay; Chemistry Fourth Edition.

3- The foundation of chemistry ; Published on Oct 15, 2012