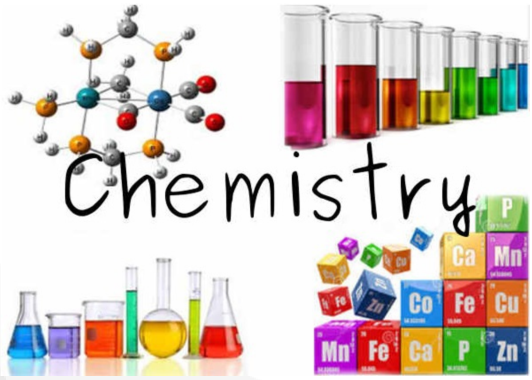
STRUCTURE OF MATTER



STRUCTURE OF MATTER

MESLI Meryem

Tlemcen University

Faculty of Technology

Hydraulics Department

1.0

Janvier 2024

Table of contents

Objectives

The subject structure of matter has several objectives. Here are some of the general objectives that might be encountered:

1. Understanding the fundamental nature of matter: The module aims to provide an in-depth understanding of what matter is made of, exploring atoms, subatomic particles and the fundamental forces that govern them.

2. Explain the basic principles of quantum physics: This includes familiarisation with concepts such as the uncertainty principle, the standard model of particle physics, and wave-particle duality.

3. Understand the structure of the atom: This involves studying the configuration of electrons around the nucleus, energy levels, atomic orbitals and packing rules.

4. Analysing the nuclear structure: This involves studying the protons and neutrons inside the atomic nucleus, and the forces that hold them together.

5. Exploring fundamental interactions: This involves understanding the four fundamental forces of nature (gravitation, electromagnetism, strong interaction and weak interaction) and their role in the functioning of the universe at microscopic scales.

6. Applying concepts to observable phenomena: Students are often encouraged to use their knowledge to explain macroscopic phenomena, such as electrical conductivity, electromagnetic radiation and the effects of electromagnetic fields.

Together, these aims are intended to provide students with a solid grounding in fundamental physics and chemistry, as well as the tools needed to understand and analyse various observable phenomena in the world around them.

to study a course “the structure of matter”, it is generally necessary to have a solid grounding in physics and a knowledge of mathematics, particularly differential and integral calculus and prior knowledge of fundamental chemistry.

prerequisite tests :

test 1 : calculate the molar mass of the following molecule, glucose, water and carbon dioxyde

I Fundamental Concept

Chemistry is the study of matter and the transformations it can undergo. All that has a mass and occupies a space is matter.

the true impact of chemistry extends much farther than the products we use in daily life.

The most profound questions about health, climate change, even the origin of life, ultimately have chemical answers.

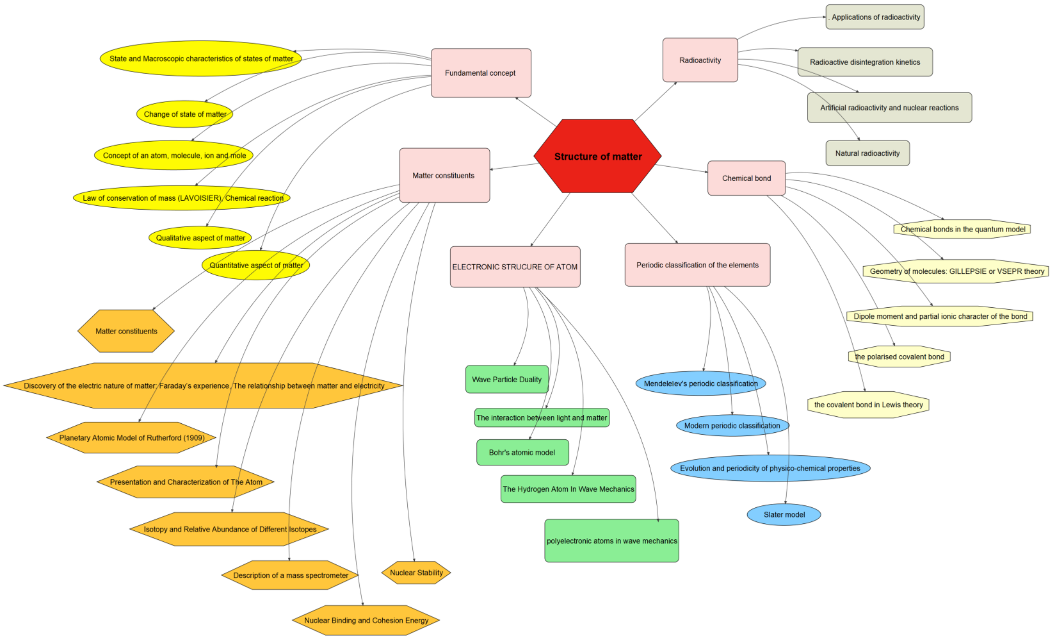
No matter what your reason for studying chemistry, this course will help you develop two mental skills. The first, common to all science courses, is the ability to solve problems systematically. The second is specific to chemistry, for as you comprehend its ideas, you begin to view a hidden reality filled with incredibly minute particles moving at fantastic speeds, colliding billions of times a second, and interacting in ways that determine how all the matter inside and outside of you behaves. This chapter holds the keys to enter this world.

to study a course “the structure of matter”, it is generally necessary to have a solid grounding in physics and a knowledge of mathematics, particularly differential and integral calculus,

A good place to begin our exploration of chemistry is to define it and a few central concepts. Chemistry is the study of matter and its properties, the changes that matter undergoes, and the energy associated with those changes.

Definition of the matter

Matter is the “stuff ” of the universe: air, glass, planets, students—anything that has mass and volume.



1. State and Macroscopic characteristics of states of matter

1.1. State of matter

Matter comes in different forms called states of matter

The three states of matter are: solid, liquid and gas

Each state of matter has its own characteristics, such as:

- ***The solid*** state keeps a constant shape and volume, cannot be compressed and cannot drain.

- ***The liquid*** state takes the shape of the container, keeps the volume constant, cannot be compressed, can flow

- ***The gaseous state*** takes the shape and volume of the container, can be compressed and can flow

1.2. Properties of matter

• ***Physical properties***: a characteristic that can be observed or measured without changing the nature of matter.

Examples: mass, temperature, color

• ***Chemical properties***: Any property that describes how a material reacts with another material, forming a new material.

Example: the reaction of zinc with hydrochloric acid gives hydrogen gas.

1.3. Physical quantities characteristics of matter and its states

The physical quantities can be expressed by the 5 basic units defined by the international system (SI).

2. Change of state of matter

Each status change has a specific name:[1]

• When a solid becomes liquid, it is called fusion

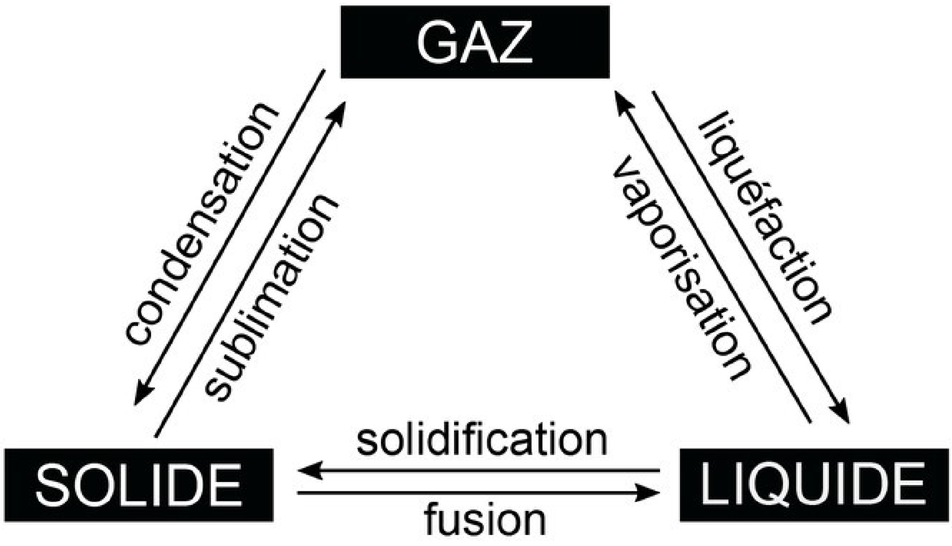
• When a liquid becomes solid, it is called solidification

• When a liquid becomes gas, it is called vaporization

• When a gas becomes liquid, it is called liquefaction

• When a gas becomes solid, it is called condensation

• When a solid becomes gas, it is called sublimation



3. Pure substance

A pure substance is a substance that contains only one kind of particles

Examples: distilled water, table salt (sodium chloride), oxygen gas

A pure substance can be simple or composed.

- Element consists of a single element type.

Examples:

H (hydrogen), Na (sodium), Fe (iron)

-compound is a substance that consists of at least two different elements in defined proportions.

Example:

water is a compound of hydrogen and oxygen, with two hydrogen atoms per oxygen atom

4. Concept of an atome, melecule, ion and mole

• ***The atom*** is the smallest part of an element that can exist.

Ex: the smallest part of the carbon element C is a carbon atom

• Atoms combine to form molecules, so a molecule is a union of atoms.

Ex: the water molecule is formed by the union of 2 hydrogen atoms H and 1 oxygen atom O

• ***The molecule*** is represented by a chemical formula

It indicates the number of atoms of each element in a molecule.

Example:

the water molecule has the molecular formula H2O: each water molecule contains two hydrogen atoms H and one oxygen atom O

• ***An ion*** is an atom or molecule charged positively (cation) or negatively (anion)

Ex: cation sodium (Na+); anion chloride (Cl-); cation ammonium (NH4+)

• ***A mole*** is the number of particles (atoms, molecules or ions) equal to the number of carbon atoms in 12 g of C-12

This number is: 6.022 x 1023. called Avogadro number

1 mol of particles = 6.022 x 1023 particles

• ***Molar mass***

The molar mass of an element is the mass, in grams, of 6.022 x 1023 atoms of this element: It is the mass of 1 mol of this element

This mass appears on the periodic table [2]

5. Atomic mass unit

In chemistry, a unit for expressing masses of atoms, molecules, or subatomic particles. An

atomic mass unit is equal to 1/12 the mass of a single atom of carbon-12.

The mass of an atom consists of the mass of the nucleus plus that of the electrons, so the

atomic mass unit is not exactly the same as the mass of the proton or neutron. Atomic mass

units are also called daltons (Da), for chemist John Dalton∗.

Atomic mass is the mass of an atom expressed in units of atomic mass (uma, u).

By definition:

1 atom ^{12}C has a mass of 12 uma.

1 uma = 1.6605×10^{-27} kg

Examples:

– 1 atom of carbon = 12.01 uma

– 1 atom of oxygen = 16.00 uma

– 1 molecule O2 = 2(16.00 uma) = 32.00 uma

Definition

1 atom 12C has a mass of 12 uma.

1 uma = 1.6605×10^{-27} kg

6. Density

Density of a solution (ρ) is the mass of a specified volume of the solution (g/cm^3 = g/mL = kg/dm^3 = kg/L); it is often labelled on a bottle containing the solution.

Example: ρ = 1,8 g/mL means that 1 mL of the solution weights 1,8 g).

Molar mass

This is the volume occupied by one mole of substance in the gaseous state. Under Normal Pressure and Temperature conditions (P= 1 Atm, T= 0°C=273 K), one mole of gaseous substance occupies a volume of 22.4 L

7. Law of conservation of mass (LAVOISIER), Chemical reaction

the law of conservation of mass was discovered by Antoine Lavoisier in 1789 (French chemist), this law states that in chemical reactions the total mass of reagents equals the total mass of products. No atoms are lost or created during chemical reactions, only rearranged, maintaining a consistent mass throughout.

You can write an equation balance of a chemical reaction:

Chemical reaction

Reactant →formed products

- In a chemical reaction, the elements ***retain***

- In a chemical reaction, the mass of the l***ost*** reagents is ***equal*** to the mass of the products formed (Lavoisier's Law)

Nothing is lost, nothing is created, everything is transformed [1]

8. Qualitative aspect of matter

8.1. Mixtures

A mixture has at least two different pure substance, so at least two types of particles

Examples: seawater, brass (copper and zinc), plants, medicines, perfume, etc.

We have two types of mixtures:

a) Heterogeneous mixtures:

It is a mixture for which we can distinguish at least 2 constituents (with the naked eye or under the microscope)

Examples: most of the rocks, mixtures of sugar and sand, milk ...

b) Homogeneous mixtures:

it is a mixture for which one does not distinguish the different constituents with the naked eye (and even with a microscope)

Example: sugar water, salt water, syrup, honey, etc.

8.2. Solutions

Solutions are homogeneous mixtures of at least two substances.

- ***The solvent*** is the substance present in large quantities
- ***The solute*** is the substance present in small quantities and which is dissolved in the solvent.

When the solvent is water we speak of aqueous solution

9. Quantitative aspect of matter

9.1. Molarity

Molarity (M, C) is one of the most commonly used measures of concentration in chemistry it is defined as the ratio between the number of moles of solute, n, and the volume of the solution, V, expressed in litres of solution (mol/l)[3]

Example:

Calculate the molarity of a prepared solution by dissolving 15 g of NaOH in enough water to make 250 mL of solution. M(NaOH) = 40g/mol

$$C = \frac{n}{V}$$

9.2. Molality (m)

Molality (m) does not contain volume in its ratio; it is the number of moles of solute dissolved

in 1000 g (1 kg) of solvent:

$m = n_{solute} / W_{solvent}$

9.3. Mass concentration

The mass concentration of a solution is equal to the mass m of solute divided by

the volume V of the solution. of the solution, expressed in $g.L^{-1}$

$C_m = m/V$

where:

C_m = mass concentration in $g.L^{-1}$

m mass of the solute in g and V the solution volume in L.

9.4. Normality

Normality is mainly used as a measure of reactive species in a solution and during titration

reactions or particularly in situations involving acid-base chemistry.

As per the standard definition, normality is described as the number of gram or mole

equivalents of solute present in one litre of a solution. When we say equivalent, it is the

number of moles of reactive units in a compound.

9.5. Dilution

Solutions used in the laboratory are often purchased in a concentrated form. From these commercial solutions, it is possible to prepare solutions of lower concentrations by performing dilution, that is, by adding solvent. To make a dilution, it is necessary to know how to pass the concentration of a solution from an initial value C_i to the desired value C_f.

After dilution, the quantity of substance is the same. We can write :

Before dilution: $n1 = C1.\ V1$

After dilution: $n2 = C2.\ V2$

Number of moles is the same: $n1 = n2 \Rightarrow C1.\ V1 = C2.\ V2$

$$n = C \times V$$

$$Ci \times Vi = Cf \times Vf$$

9.6. Exercises

Exercise 1

Which of the following samples contains the most iron? 0.2 moles Fe2(SO4)3 ,20g iron

iron, 0.3 atom-gram of iron 2.5x1023 atoms of iron

Data: MFe=56g.mol-1 MS =32g.mol-1

Avogadro number N =6.023. 1023

Solution

Reminder: In one mole, there are N particles (atoms or molecules)

*0.2 moles of Fe2(SO4)3 corresponds to 0.4 moles of atoms (or gram-atom) of iron.

*20g of iron corresponds to n= m/MFe = 20/56 = 0.357 moles of iron atoms, 0.3 atom-grams of iron or 0.3 moles of iron atoms.

*2.5x1023 iron atoms corresponds to n = number of atoms, N= 0.415 moles of iron atoms.

This last sample contains the most iron.

10. exercises

Exercise 1

1. How many moles are there in: 4 g of NaOH ; 30 mL H_2SO_4 (d= 1,83); 100 µg of $KMnO_4$; 2,75 10^{32} atoms of iron (Fe).

2. Which sample is the most iron-rich: 2 g of $Fe_2(SO_4)_3$ and 5,30 x 10^{21} atoms of iron.

3. Which sample contains the least moles of atoms : [25 g of carbone or 2,49 10^{22} atoms of Au (or)]

4. Calculate in g and in Kg the corresponding mass at 1 u.m.a.

Data : molair mass (g/mol): C (12) ; Na (23) ; O (16) ; S (32) ; K (39) ; Mn(55) ; Fe(56) ; Cl (35,5)

Exercise 2

a-Calculate the molarity of solution A preparated by dissolving 4,2 g of NaOH in distilled water to obtain 350 ml of this solution.

b-what is the volume of distilled water added to the solution A to obtain solution B at 0,25 M.

II Exercises

Exercise 1

We assume that the mass of the phosphorus atom 15P is equal to the sum of the masses of the particles that make it up.

1) What is the mass of the nucleus of a phosphorus atom?

2) What is the mass of the electron cloud of a phosphorus atom?

3) What is the mass of a phosphorus atom? What conclusion do you draw?

Exercise 2

The mass of all the electrons in the iron atom is $2{,}366.10^{-29}$ kg.

1- Knowing that one electron has a mass of $9{,}1.10^{-31}$ kg, how many electrons does an iron atom have?

2- What is the number of positive charges carried by the nucleus of the iron atom?

3- Deduce the atomic number of the iron atom. The mass of an iron atom is $9{,}3.10^{-26}$ kg.

4- Calculate the number of iron atoms that make up an iron nail of 2,5 g.