PEOPLE'S DEMOCRATIC REPUBLIC OF ALGERIA

MINISTRY OF HIGHER EDUCATION AND SCIENTIFIC RESEARCH

University Ferhat Abbas Setif-1

Faculty of Natural and Life Sciences

General and Organic Chemistry

- ✓ Fundamental notions
- ✓ Radioactivity
- ✓ Electron configuration
- ✓ Chemical bonds
- ✓ Organic chemistry

Common base L1: Natural and Life Sciences

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Foreword

Chemistry is the science that studies the composition, structure and properties of matter, as well as the transformations that it can undergo. It is at the heart of understanding many aspects of nature and life. The size of the compounds studied in chemistry varies from reactions between simple atoms to molecular structures of several tens of thousands of atoms (DNA, proteins, crystals, etc.). Studying the world at the molecular level allows us to better understand the world at the human level. Indeed, chemistry is part of everyday life. Using light, plants produce the nutrients they need by capturing gases dissolved in the air (mostly nitrogen) and drawing water and mineral salts from the soil. They do chemistry. When we heat food, when we cook an egg, pasta, etc. We also do chemistry. All beauty products, perfumes and creams have a chemical origin. They are created or synthesized from atoms and molecules that react with each other. Either by imitating and reproducing what happens in nature, or by creating new compounds.

Chapter I focuses mainly on the fundamental notions of general chemistry with a reminder of the states of matter, atoms, molecules and solutions.

Chapter II deals with radioactivity and nuclear reactions.

Chapter III deals with the electronic configuration of atoms using the main laws of filling quantum boxes and involving the four quantum numbers.

Chapter IV deals with the principle of the periodic classification of the elements, the rules for filling atomic orbitals, the evolution and periodicity of the physicochemical properties of the elements.

Chapter V is devoted to chemical bonding: ionic bonding, covalent bonding, Lewis's structure and predicting the geometry of molecules using the VSEPR method.

The last chapter is reserved for the basic notions of organic chemistry.

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CHAPTER I: Fundamental notions of Chemistry I

I.1. Introduction

Matter appears in the form of solid, liquid or gas, it is capable of passing from one state to another under very specific temperature and pressure conditions.

a) **Solid:** It is rigid, the shape and volume of the solid matter are almost constant and the arrangement of its particles follows a well-defined order, we cite as an example: salts.

b) **Liquid:** The particles in liquid solutions are more agitated than in the solid state, it has a fixed volume but a variable shape, it takes the shape of its container.

c) **Gas:** In this state the atoms and molecules are excessively agitated and therefore very disordered. They occupy a very large volume; the gas takes the shape of its container.

In fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas).

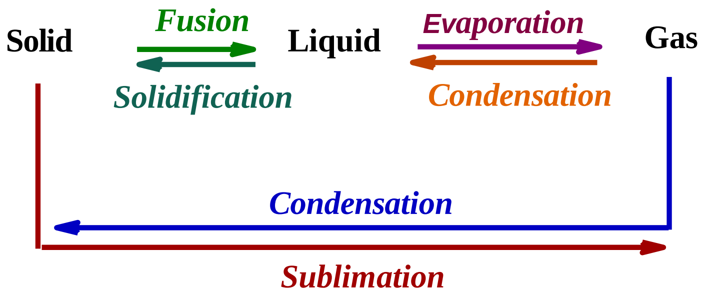


Figure I.1: Phase changes.

I.2. Concepts of atoms, molecules, moles and Avogadro’s number

Matter is made up of elementary grains called atoms. The atom is an infinitely small quantity of matter; the order of magnitude of its mass is 10^{-26} kg. The size of the atom is a few Angstrom (A°). Atoms combine to form molecules. A molecule is therefore an union of atoms.

The Avogadro number of atoms or molecules corresponds to 1 mole of atoms or molecules. NA= 6.023 10^{23}, 1 mole of O_2 equal to 16 g.

I.2.1. Atomic mass unit (amu)

By definition, 1uma corresponds to 1/12 of the mass of a carbon atom.

1uma = 1/12 (mass of a Carbon atom) = 1/12 (1x12g)/ NA = 1/NA= 1/6.023 10^{23} = 1.66 10^{-27} Kg.

I.2.2. Atomic, molecular molar mass

Atomic molar mass is the mass of 1 mole of atoms or NA of atoms.

Molecular molar mass (M) is the mass of 1 mole of molecules or NA of molecules.

I.2.3. Molar volume

We define the molar volume as being the volume of one mole of substance, this expression is used mainly for gases.

I.3. Qualitative aspect of the material

Matter appears in a pure body (simple or compound) or in a mixture state.

I.3.1. simple pure body

The material that composes it is a single chemical species distinguished by physicochemical properties such as density, melting and boiling temperatures, odor, color, etc. Like H_2, O_2, N_2…

I.3.1.1. Solute

Is a chemical species (molecular or ionic) dissolved in a solvent.

I.3.1.2. Solvent

It is the liquid in which the solute is always dissolved in a quantity much greater than the solute(s). The solvent can be distilled water, ethanol, etc.

I.3.1.3. Solution

This is the name for the homogeneous mixture of solute and solvent. When the solvent is water, the solution is called aqueous. The dissolved species is characterized by different types of concentration.

I.4. Concentration of a dissolved species

I.4.1. Molarity (mol/L).

It is equal to the quotient of the quantity of material of A dissolved by the volume V of solution. CM = n moles of solute / V solution

I.4.2. Mass concentration (g/L)

It is equal to the ratio of the mass m (in g) of dissolved solute by the volume V (in l) of solution Cm = m solute (g) / V solution (l)

I.4.3. Mole fraction

Is the ratio of the quantity of material of compound (i) in moles contained in a certain volume of solution divided by the sum of the quantities of material of all the constituents i,j… present in this volume of solution. $X_i = n_i \text{ (mole)} / \Sigma n_{i,j} \text{ (mole)}$.

Example:

We have an aqueous solution of ethanol, we can calculate the mole fraction of each constituent of this solution. X_{ethanol} = n ethanol / (n ethanol + n H_2O).

I.4.4. Molality

It corresponds to the quantity of material of X for 1Kg of solvent. This concentration unit is very rarely used. It is denoted by C_{molale}.

I.4.5. Normality

The normality of an acidic or basic solution in water is defined as respectively the number of moles of (Z) H^+ or OH^- ions likely to be released by one liter of solution.

Similarly, the oxidation-reduction normality of a solution corresponds to the number of moles of electrons (Z) likely to be released by one liter of solution.

It is denoted by N and is calculated according to the following formula:

N = number of gram equivalents / V solution (L)

$N = CM \times Z$

Z: depends on the nature of the solution.

CM: Molar concentration of the solution

- ✓ Acid solution: CH_3COOH, Z=1, H_2SO_4 Z=2.
- ✓ Basic solution: NaOH, Z=1
- ✓ Salts: $FeSO_4 \leftrightarrow Fe^{2+} + SO_4^{2-}$ $Z = |+2 \times 1|$ or $|-2 \times 1|$

Redox normality: Z is the number of electrons exchanged between the reactive ions in the solution.

I.4.6. Mass fraction (purity)

It is the ratio of the mass of a compound (solute) X, contained in a certain volume of solution divided by the mass of the solution. For example, a 10% solution of X contains 10g of X per 100g of solution. % mass = (m solute / m solution) × 100

I.4.7. Volume percentage

It is the ratio of the volume of compound (solute) X divided by the volume of the solution.

% Volume= (V solute ⁄ solution) ×100

I.4.8. Volumic mass ρ

Is the mass per unit volume of a substance, i.e. the mass m of a substance of volume V.

$P = m_{solution} / V_{solution}$

With:

m: is the mass in Kg, g.

V: is the volume in m^3, l or cm^3

Then ρ is expressed in kg/m^3, g/l or g/cm^3.

I.4.9. Density

The density of a substance is defined relative to that of water.

$d_{substance} = \rho_{substance} / \rho_{solvent}$

The two densities being expressed in the same unit, it follows that the density
is a dimensionless number.
For gases, the density is calculated with respect to the density of air.

$d_{gas} = \rho_{gas} / \rho_{air}$

The density of air is low (1.2 kg/m^3 at sea level at 20°C) compared to that of water (1000 kg/m^3 at 20°C).

I.5. Dilution of an aqueous solution

The dilution of an aqueous solution consists in reducing its concentration by adding solvent (water). The initial solution of higher concentration is called stock solution. The final solution of lower concentration is called the daughter solution (dilute solution). During a dilution, there is conservation of the quantity of matter of solute so that one can write:

$$n_i = n_f \; C_i \times V_i = C_f \times V_f$$

With:

n_i: Quantity of initial material relative to the stock solution.

n_f: Quantity of final material relative to the daughter solution (diluted).

V: Volume.

C: Concentration.

Example:

We have 100 ml of aqueous solution of diode I_2 with a concentration of 5.10^{-2} mol. l^{-1}. What volume of stock solution should be used to produce 50 mL of diluted iodine solution with a concentration of 1.10^{-2} mol. l^{-1}?

I.5.1. Dilution in biology

Dilutions are carried out in many cases, for example:

a) Dosage reaction for a protein: albumin using a standard range. The standard range is a dilution series to be carried out from a stock albumin solution (reference) of known concentration.

b) Blood count: We take a small volume of fresh blood, dilute it 1/100 in physiological water. We take some of this dilution and count the blood cells under a microscope.

https://youtu.be/4VltXjR64SU

https://study.com/academy/topic/fundamental-concepts-in-chemistry.html

Chapter II: Radioactivity

II.1. Introduction

Matter is composed of indivisible atoms. An atom is an extremely small particle of matter that retains its identity during chemical reactions.

The atom is composed of three elementary particles of very small size which are:

- Electrons (discovered by Crooks in 1879).
- Protons (discovered by Rutherford in 1918).
- Neutrons (discovered by Chadwick in 1932).

These three particles constitute all the matter in the universe and therefore all the elements of nature.

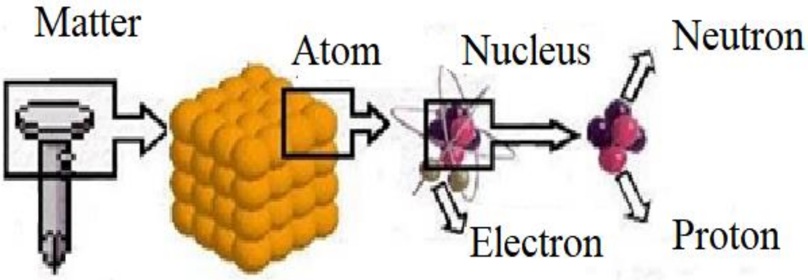
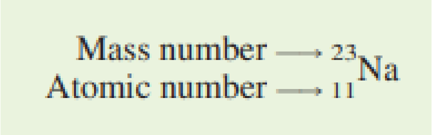


Figure II.1: Main constituents of matter.

The shorthand notation for any nuclide consists of the symbol of the element with the atomic number written as a subscript on the left and the mass number as a superscript on the left. You write the nuclide symbol for the naturally occurring sodium nuclide as follows:



Atoms can have the same Z and different A's. They are said to be isotopes of the same element.

Table II.1: Some natural isotopes.

Elements	${}^{A}_{3}Li$	${}^{A}_{8}O$	${}^{A}_{17}Cl$	${}^{A}_{16}S$
A	6 7	16 17 18	35 37	32 33 34 36

II.2. Mass defect and cohesion energy

II.2.1. Mass defect

We call the mass defect of a nucleus the difference between the total mass of the A separated nucleons (Z protons + N protons), at rest and the mass of the nucleus formed at rest. The mass m of a nucleus is always less than the sum of the masses of its isolated nucleons.

$$\boldsymbol{\Delta m = m}_{\text{theoretical}} - \boldsymbol{m}_{\text{calculate}}$$

When there is a nuclear reaction:

$$\Delta m = \sum m_{\text{reactants}} - \sum m_{\text{products}}$$

Example the following reaction:

$$2p + 2n \longrightarrow {}^{4}_{2}He + E$$

$m_{\text{theoretical}} = Z.m_p + N.m_n = 2 \times 1{,}0073 + 2 \times 1{,}0087 = 4{,}03200$ u.m.a

$m_{\text{calculates}} = m_{\text{exp}} = 4{,}00150$ u.m.a

$\Delta m = 0{,}03050$ u.m.a

II.2.2. Cohesion energy of nuclei

The formation of a nucleus from these different constituents occurs with very significant energy absorption. EINSTEIN’s principle of mass-energy equivalence predicts that a quantity of energy will be released.

$$\Delta \mathbf{E}\ (\text{MeV}) = \Delta \mathbf{m} \times \mathbf{C}^{2}$$

C: speed of light in a vacuum = 3.10^{8} m/s

By definition, cohesion energy (E_{ch}) is the energy that a nucleus would need to bring to dissociate it into its nucleons. It is a positive energy:

$${}^{A}_{Z}X + Ech \rightarrow Z + N$$

The most commonly used unit of energy is eV (KeV = 103 eV, MeV = 106 eV)

Deuterium ${}^{2}D$ is an isotope of hydrogen. We give:

m_n =1,008665 u.m.a, m_p =1,007277 u.m.a and m_{noyau} = 2,014102 u.m.a

Calculate the energy of this stable isotope, in joules and MeV?

Δm =0.00184 a.u.m = $3.055.10^{-30}$ Kg

We transform the u.m.a into Kg

$E = \Delta m \times C^{2}$

Ech = 2.710^{-13}J = 1.685 MeV

II.2.3. Cohesion energy per nucleon

In order to compare the stability of nuclei, it is preferable to think in terms of cohesion energy per nucleon defined as: $\frac{E_{ch}}{A}$ is expressed in MeV/nucleon. The greater the ratio, the more stable the nuclei.

II.3. Radioactivity

The phenomenon of radioactivity was discovered by Antoine Henri Becquerel in 1896. He discovered that photographic plates develop bright spots when exposed to uranium minerals, and he concluded that the minerals give off some sort of radiation.

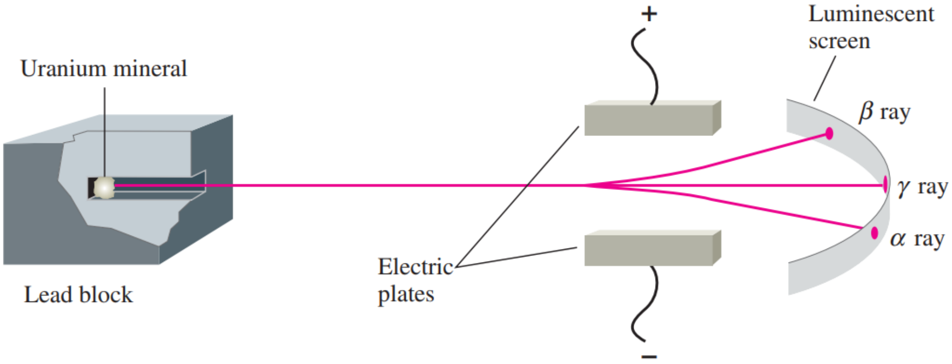


Figure II.2: Separation of the radiation from a radioactive material.

We will look at two types of nuclear reactions. One type is radioactive decay, the process in which a nucleus spontaneously disintegrates, giving off radiation. The second type of nuclear reaction is a nuclear bombardment reaction, a nuclear reaction in which a nucleus is bombarded, or struck, by another nucleus or by a nuclear particle.

II.3.1. Radioactive decay (Natural radioactive)

An unstable or radioactive nucleus is a nucleus that spontaneously (naturally) decays into a different nucleus and emits three types of radiation (α, β or γ).

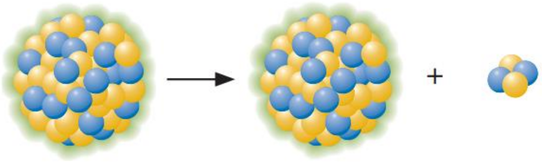
II.3.1.1. Alpha rays (α)

Emission of α $^{4}_{2}He$ nucleus, or alpha particle, from an unstable nucleus.

α radiation has a strong ionization force and a low penetrating power: they are particles easily stopped by a few centimeters of air or a sheet of paper.

$$^{A}_{Z}X \rightarrow {}^{A-4}_{Z-2}Y + {}^{4}_{2}He$$

An example is the radioactive decay of radium-226.



$^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$

II.3.1.2. Beta emission (β or β^-)

Emission of a high-speed electron from an unstable nucleus. Beta emission is equivalent to the conversion of a neutron to a proton. When $\frac{A-Z}{Z} \geq 1.5$, the nucleus contains an excess of neutrons, such nuclei will seek to stabilize by increasing Z and decreasing N by transforming a neutron into a proton.

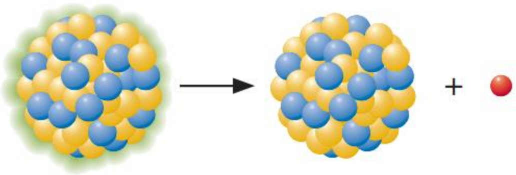
$^{1}_{0}n$ ● $\longrightarrow$ $^{1}_{1}p$ ● + $^{0}_{-1}e$ ●

II.3.1.3. Positron emission (β^+)

Emission of a positron from an unstable nucleus. A positron, denoted in nuclear equations as $^{0}_{1}e$, is a particle identical to an electron in mass but having a positive instead of a negative charge. Positron emission is equivalent to the conversion of a proton to a neutron.

$^{1}_{1}p$ ● $\longrightarrow$ $^{1}_{0}n$ ● + $^{0}_{1}e$ ●

The radioactive decay of technetium-95 is an example of positron emission.

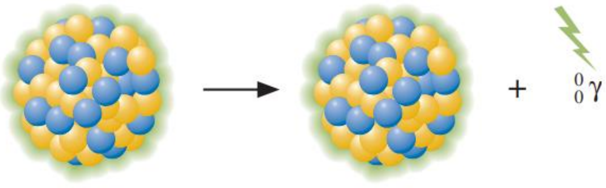


$^{95}_{43}Tc \longrightarrow {}^{95}_{42}Mo + {}^{0}_{1}e$

The product nucleus has an atomic number that is one less than that of the original nucleus. The mass number remains the same.

II.3.1.4. Gamma rays (γ)

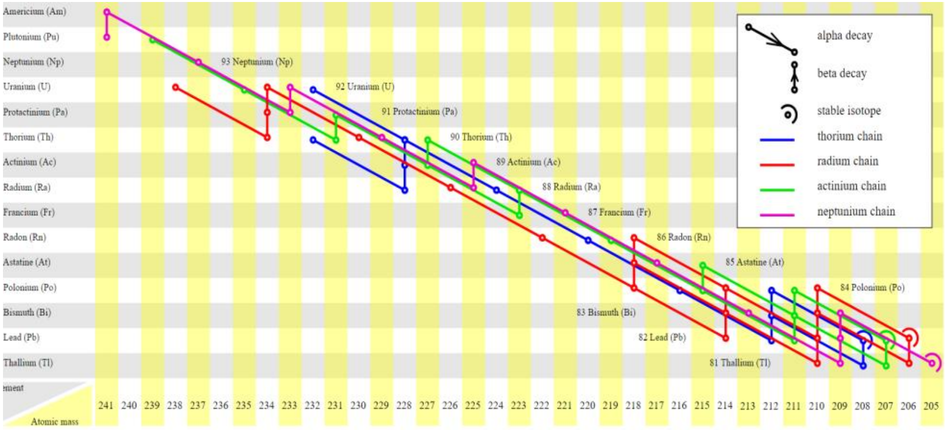
Emission from an excited nucleus of a gamma photon, corresponding to radiation with a wavelength of about 10^{-12} m. In many cases, radioactive decay results in a product nucleus that is in an excited state.



$$^{99m}_{43}\text{Tc} \longrightarrow {}^{99}_{43}\text{Tc} + {}^{0}_{0}\gamma$$

The product nucleus is simply a lower-energy state of the original nucleus, so there is no change of atomic number or mass number. γ rays are more penetrating than alpha and beta particles, but are less ionizing.

Natural radioactive elements belong to four fundamental families:

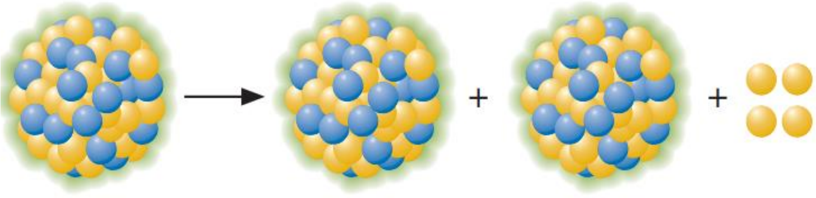


II.3.2. Artificial radioactivity

It is a nuclear reaction which consists of the bombardment of a stable or radioactive nucleus by neutrons, protons or α particles with the aim of creating a new artificial radioactive element which does not exist in nature.

II.3.2.1. Spontaneous fission

The spontaneous decay of an unstable nucleus in which a heavy nucleus of mass number greater than 89 splits into lighter nuclei and energy is released. For example, an uranium-236 atom can spontaneously undergo the following nuclear reaction:



$$^{236}_{92}U \longrightarrow ^{96}_{39}Y + ^{136}_{53}I + 4^{1}_{0}n$$

Atoms with very high mass numbers (A>200), when bombarded by neutrons can undergo breakage leading to lighter atoms and neutrons. The neutrons emitted can in turn cause the fission of neighboring atoms, resulting in an explosive chain reaction. During this process there is a loss of mass and a significant release of energy. This type of chain reaction when uncontrolled is the basis of the atomic bomb.

$$^{235}_{92}U + ^{1}_{0}n \rightarrow \qquad ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n$$

II.3.2.2. Nuclear fusion

Light atoms will seek to stabilize by fusion reaction. During this type of reaction, two light nuclei will fuse to produce a heavier atom and various particles. During this process there will be a loss of mass and a significant release of energy. This is the reaction used in H-bombs.

$$^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n$$

II.4. Laws of radioactivity

A^* ⟶ B (stable)

t = 0 N_0 (Number of radioactive nuclei at t=0)

t > 0 N_t (Number of radioactive nuclei remaining at time t)

II.4.1. Decay rate

$$-\frac{dN}{dt} = \lambda N$$

λ: radioactive constant

$$\frac{dN}{N} = -\lambda dt$$

By integration:

$$\int_{N0}^{Nt} \frac{dN}{N} = -\lambda \int_0^t dt \qquad \ln N_t - \ln N_0 = -\lambda t \qquad \ln\frac{Nt}{N0} = -\lambda t \qquad N_t = N_0 e^{-\lambda t}$$

II.4.2. Half-life of a radioactive nucleus $t_{1/2}$

We define the half-life of a radioactive nucleus as the time it takes for one-half of the nuclei in a sample to decay.

$$t = t_{1/2} \qquad \frac{N_0}{2} = N_0 e^{-\lambda t_{1/2}}$$

$$t_{1/2} = \frac{ln2}{\lambda}$$

We found that the decay constant λ for $^{99}_{43}Tc$ was equals to $1.10^{-13}s^{-1}$. What is the half-life of this isotope in years?

II.4.3. Radioactive activity

The average activity A of a radioactive sample is the average number of decays that occur per second:

$$A(t) = -\frac{dN}{dt} \qquad A_t = A_0 e^{-\lambda t}$$

A_t (activity) at time t and A_0 (initial activity).

Unit of A is the Becquerel (1975) we have: 1Bq = 1 dps (disintegration per second) Other unit: the Curie activity of 1 g of radium: 1Ci = $3.7\ 10^{10}$ Bq or dps.

https://www.youtube.com/watch?v=f6NOTo3YAts

Chapter III: Electron configuration

III.1 Quantum numbers and atomic orbitals

The allowed values and general meaning of each of the four quantum numbers of an electron in an atom are as follows:

III.1.1 Principal Quantum Number (n)

This quantum number is the one on which the energy of an electron in an atom principally depends; it can have any positive value: 1, 2, 3, and so on.

Orbitals of the same quantum state n are said to belong to the same shell. Shells are sometimes designated by the following letters:

n	1	2	3	4	5	6	7
Letter	K	L	M	N	O	P	Q

III.1.2 Angular Momentum Quantum Number (l)

Also Called Azimuthal Quantum Number, this quantum number distinguishes orbitals of given n having different shapes; it can have any integer value from 0 to n - 1. Within each shell of quantum number n, there are n different kinds of orbitals, each with a distinctive shape denoted by an l quantum number.

$$0 \leq l \leq n-1$$

Orbitals of the same n but different l are said to belong to different subshells of a given shell. The different subshells are usually denoted by letters as follows:

l	0	1	2	3	4
subshells	s	p	d	f	g

III.1.3 Magnetic Quantum Number (m)

This number defines the number of orientations in space that the electron can take when subjected to the action of a magnetic field. It characterizes the quantum box occupied by the electron. For l = 0 (s subshell), the allowed m quantum number is 0 only; there is only one orbital in the s subshell.

Note that there are 2l +1 orbitals in each subshell of quantum number l.

$$-l \leq m \leq +l$$

To symbolize the different shapes of the orbital, we use a quantum box represented by a rectangle. There are as many rectangles as there are possible values of m.

☐

III.1.4 Spin Quantum Number (m_s)

This quantum number refers to the two possible orientations of the spin axis of an electron; possible values are +1/2 and -1/2. An electron acts as though it were spinning on its axis like the earth.

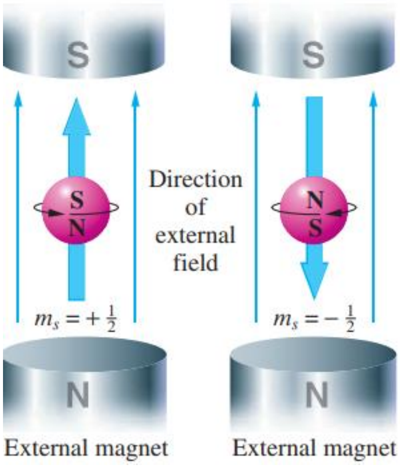


Figure III.1: A representation of electron spin.

We represent this spin number symbolically by:

An upward arrow (↑): an electron in the spin state +1/2.

A downward arrow (↓): an electron in the spin state – 1/2.

The S orbital:

n = 1, l = 0, m = 0, S We represent this orbital by a rectangular box.

(s = +1/2 or -1/2) so the S orbital contains at most 2e-.

Shell K($2e^-$).

☐

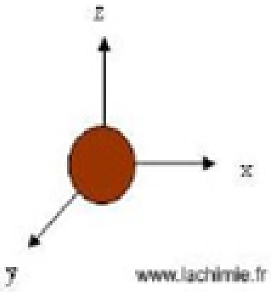


Figure III.2: Graphical representation of the O.A nS.

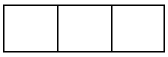
The p orbital:

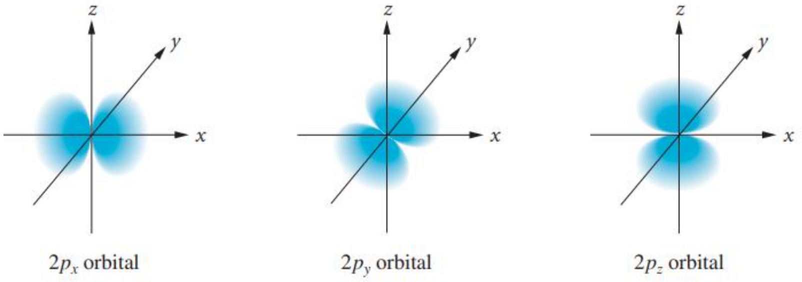
n = 2,

l = (0), m = (0) S (max 2e^{-})

l = (1), m = (-1, 0, 1) P (max 6e^{-})

Electron distribution in the 2px orbital. Note that it consists of two lobes oriented along the x-axis.



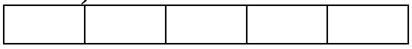


$2p_x$ orbital $2p_y$ orbital $2p_z$ orbital

Figure III.3: The 2p orbitals.

The d orbital:

l = 2, m = (-2, -1, 0, 1,2) d (max 10 e^{-})



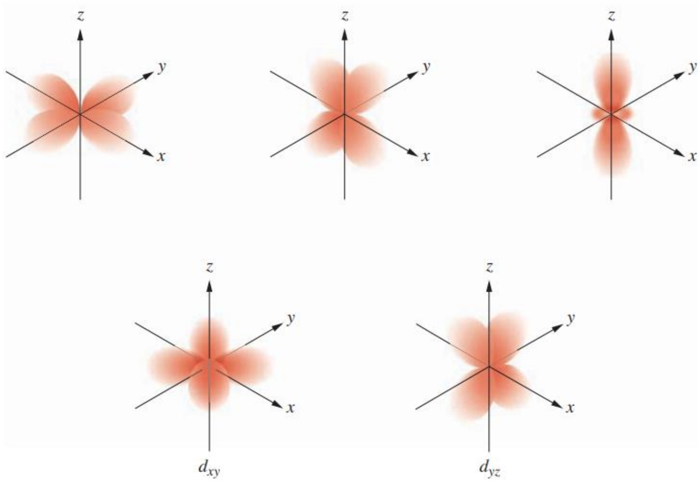


Figure III.4: The five 3d orbitals.

The f orbital:

l = 3, m = (-3, -2, -1, 0, 1, 2,3) f (max 14 e^{-}).

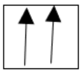
III.2 Electron Configurations and Orbital Diagrams

An electron configuration of an atom is a particular distribution of electrons among available subshells. The notation for a configuration lists the subshell symbols, one after the other, with a superscript giving the number of electrons in that subshell.

we use an orbital diagram to show how the orbitals of a subshell are occupied by electrons. It is called an orbital diagram. An orbital is represented by a rectangle.

III.2.1 Pauli Exclusion Principle

The Pauli exclusion principle, which summarizes experimental observations, states that no two electrons in an atom can have the same four quantum numbers. If one electron in an atom has the quantum numbers $n = 1$, $l = 0$, $m = 0$ and $m_s = + 1/2$, no other electron can have these same quantum numbers. In other words, you cannot place two electrons with the same value of m_s in a 1s orbital.

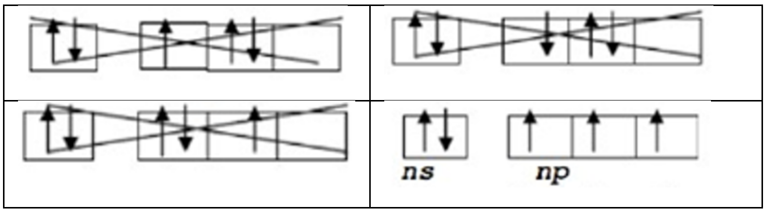


III.2.2 Principle of stability

The ground state of the atom represents the most stable state of the electron, so the electrons will start by occupying the lowest energy levels at the limit of vacant places.

III.2.3 Hund's Rule

Hund's rule states that the lowest-energy arrangement of electrons in a subshell is obtained by putting electrons into separate orbitals of the subshell with the same spin before pairing electrons.



Remark: $2n^2$ is the maximum number of electrons that can be distributed over all the subshells of the n shell.

III.2.4 Klechkowski rule

A scheme used to reproduce the electron configurations of the ground states of atoms by successively filling subshells with electrons in a specific order (the building-up order). Following this principle, you obtain the electron configuration of an atom by successively filling subshells in the following order:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f.

This order reproduces the experimentally determined electron configurations (with some exceptions, which we will discuss later). You need not memorize this order. As you will see, you can very easily obtain it from the periodic table. (You can also reproduce this order from the diagram shown in Figure III.5.

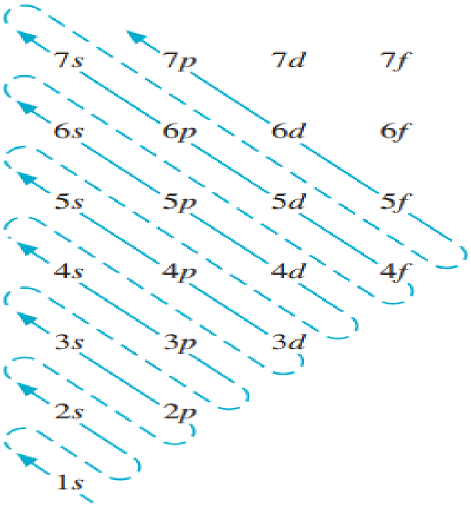


Figure III.5: Building-up order.

There are some exceptions, however, and chromium (Z=24) is the first we encounter. Copper (Z=29) is another exception to the building-up principle,

Cr (Z=24): $1s^2$ /$2s^2$ $2p^6$ /$3s^2$ $3p^6$ /$4s^2$ $3d^4$ $3d^5$ $4s^1$
Cu (Z=29): $1s^2$ /$2s^2$ $2p^6$ /$3s^2$ $3p^6$ /$4s^2$ $3d^9$ $3d^{10}4s^1$
Note: the half-filled or completely filled subshell d is more stable.

Z=9: $1S^22S^22P^5$
Z=26: $1S^22S^22P^63S^23P^64S^23d^6$
Z=37: $1S^22S^22P^63S^23P^64S^23d^{10}4P^65S^1$

III.3 Electronic configuration in condensed form

To quickly describe the electronic configuration of any element without having to write all the internal shell and subshells, we will write this configuration in condensed form:

[Noble gas configuration] + External shells

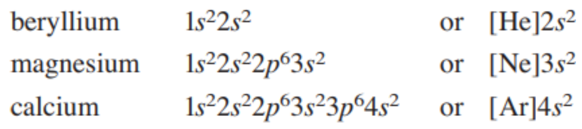
III.3.1 Noble-gas

These elements are the first members of the group called noble gases because of their relative unreactivity.

Noble gases	Symbol	Z
Helium	He	2
Neon	Ne	10
Argon	Ar	18
Krypton	Kr	36
Xenon	Xe	54
Radon	Rn	86

III.3.1 Valence electron

An electron in an atom outside the noble-gas or pseudo-noble-gas is called a valence electron. Look now at the configurations of beryllium, magnesium, and calcium, which are similar, moderately reactive elements.



beryllium	$1s^22s^2$	or	$[He]2s^2$
magnesium	$1s^22s^22p^63s^2$	or	$[Ne]3s^2$
calcium	$1s^22s^22p^63s^23p^64s^2$	or	$[Ar]4s^2$

These three elements have the same valence (2).

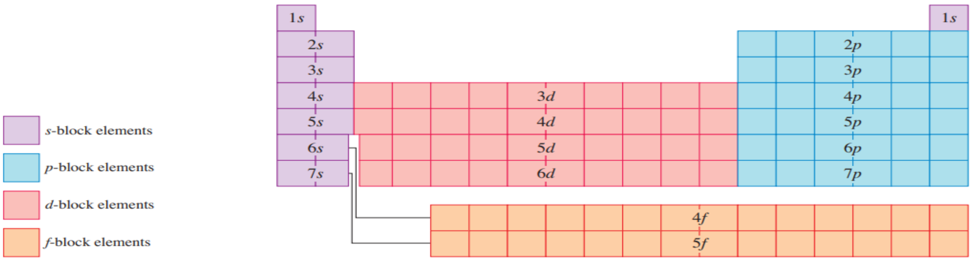
Chapter IV: Periodic Table.

IV.1 Description of Mendeleev's periodic table

IV.1.1 Periods of the periodic table

The chemical elements are classified according to increasing Z in a table with 7 rows (periods) and 18 columns (groups and subgroups). The periodic table is made up of 4 blocks: S, P, d and f. The elements of the same horizontal line constitute a period.

You read the building-up order by starting with the first period, in which the 1s subshell is being filled. In the second period, you have 2s (violet area); then, staying in the same period but jumping across, you have 2p (blue area). In the third period, you have 3s and 3p; in the fourth period, 4s (violet area), 3d (red area), and 4p (blue area).



Period	Z	Electron Configurations	Number of elements
1s	1, 2	$1s^1$ et $1s^2$	2
2s 2p	$3 \leq Z \leq 10$	$2s^{1-2}$, $2p^{1-6}$	8
3s 3p	$11 \leq Z \leq 18$	$3s^{1-2}$, $3p^{1-6}$	8
4s 3d 4p	$19 \leq Z \leq 36$	$4s^{1-2}$, $3d^{1-10}$, $4p^{1-6}$	18
5s 4d 5p	$37 \leq Z \leq 54$	$5s^{1-2}$, $4d^{1-10}$, $5p^{1-6}$	18
6s 4f 5d 6p	$55 \leq Z \leq 86$	$6s^{1-2}$, $4f^{1-14}$, $5d^{1-10}$, $6p^{1-6}$	32

The elements of the same column have the same external shell, they constitute a group or a family.

Subgroup ↗ A Contains elements whose outer layer is **ns np**.
↘ B Contains the elements whose outer layer contains subshell **d, f**

The indices I, II, III,… indicate the number of electrons in the outer shell, called (Valence electrons).

IV.1.2 Groups of the periodic table

In the following brief descriptions of the main-group elements, we will note the metallic–nonmetallic behavior of the elements, as well as the basic–acidic character of the oxides. Although elements in a given group are expected to be similar, the degree of similarity does vary among the groups.

Hydrogen ($1s^1$)

Although the electron configuration of hydrogen would seem to place the element in Group IA, its properties are quite different, and it seems best to consider this element as belonging in a group by itself. The element is a colorless gas composed of H_2 molecules.

Group IA Elements, the Alkali Metals (ns^1)

The alkali metals are soft and reactive, with the reactivities increasing as you move down the column of elements. All of the metals react with water to produce hydrogen.

Group IIA Elements, the Alkaline Earth Metals (ns^2)

The alkaline earth metals are also chemically reactive but much less so than the alkali metals. Reactivities increase going down the group.

Group VIA Elements, the Chalcogens ($ns^2\ np^4$)

These elements show the transition from nonmetal (oxygen, O, sulfur, S, and selenium, Se) to metalloid (tellurium, Te) to metal (polonium, Po).

Group VIIA Elements, the Halogens ($ns^2\ np^5$)

The halogens are reactive nonmetals with the general molecular formula X_2, where X symbolizes a halogen.

Group VIIIA Elements, the Noble Gases ($ns^2\ np^6$)

The Group VIIIA elements exist as gases consisting of uncombined atoms. These elements are known as the noble gases because of their relative unreactivity.

Group VIIIB

Iron Triad (Fe, Co, Ni). Palladium Triad (Ru, Rh, Pd). Platinum Triad (Os, Ir, Pt).

IV.1.3 Position of the element in the periodic table

The positioning of the element is done from the configuration of the external shell. Positioning of the element means its exact address in the table, namely:

a) Period takes the largest (n) of the outer shell.

b) Group the number of valence electrons.

c) Subgroup A (valence electrons s or s and p). B (valence electrons d and f)

d) Column is determined as follows.

Groupe	I		II		III		III		V		VI		VII		VIII			VIII
S-G	A	B	A	B	A	B	A	B	A	B	A	B	A	B	B			A
Column	1	11	2	12	13	3	14	4	15	5	16	6	17	7	8	9	10	18

Example: Give the electronic structure according to the rare gases and find the position of the following elements in the periodic table: ${}_{8}\text{O}$, ${}_{13}\text{Al}$, ${}_{11}\text{Na}$, ${}_{17}\text{Cl}$, ${}_{18}\text{Ar}$, ${}_{36}\text{Kr}$, ${}_{12}\text{Mg}$, ${}_{14}\text{Si}$, ${}_{35}\text{Br}$.

IV.2 Some periodic properties

The periodic law states that when the elements are arranged by atomic number (Z), their physical and chemical properties vary periodically. In this section, we will look at four physical properties of an atom: atomic radius, ionization energy, electron affinity and electronegativity.

IV.2.1 Atomic Radius

It is the distance between the center of the nucleus and the outer shell electron or it is half the distance between two atoms linked by a single bond.

1. Within each period (horizontal row), the atomic radius tends to decrease with increasing atomic number (nuclear charge). The largest atom in a period is a Group IA atom and the smallest is a noble-gas atom.

2. Within each group (vertical column), the atomic radius tends to increase with the period number.

IV.2.2 Ionization Energy

The first ionization energy (or first ionization potential) of an atom is the minimum energy needed to remove the highest-energy (that is, the outermost) electron from the neutral atom in the gaseous state.

- ✓ **Over the same period:** if Z increases then E.I increases.
- ✓ **On the same column:** if Z increases then E.I decreases.

By comparing the ionization energies; we find that: $Ei_1 < Ei_2 < Ei_3 < Ei_4 <$…

IV.2.3 Electron Affinity

This is the opposite phenomenon of ionization; The electron affinity of an atom X is the energy released when this atom captures an electron.

IV.2.4 Electronegativity (χ)

Another important property of elements is their electronegativity. Atoms can form covalent bonds to each other by sharing electrons in pairs, creating an overlap of valence orbitals.

- ✓ **Over the same period:** if Z increases then χ increases (same variation as EI).
- ✓ **On the same group:** if Z increases then χ decreases (same variation as EI).

Chapter V: Chemical bonds

V.1 Introduction

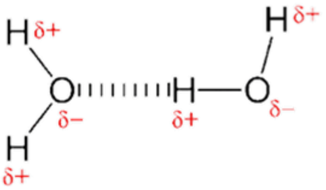
In the universe all bodies tend to acquire a state of maximum stability. Atoms are no exception: they acquire this stability by forming chemical bonds with other atoms, giving buildings called more stable molecules. The energy of a molecule is less than the sum of the energies of the isolated atoms. Noble gases whose valence layer is saturated are always found in the isolated state with very good stability.

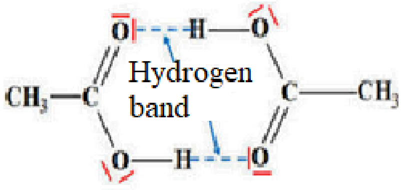
V.2 Chemical bonds

The bonds that unite atoms are called chemical bonds. These bonds are formed via valence electrons. Their nature essentially depends on the electronegativity of the atoms which is defined as the tendency of the atom to attract electrons from the bond. The difference in electronegativities between the atoms of the bond determines its nature. Depending on the dissociation energy we have two types: Strong bonds and weak bonds.

V.2.1 Hydrogen bond

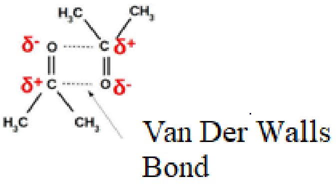
There is another type of bond called a hydrogen bridge bond. These bonds intermolecular or intramolecular forces involve an H atom bonded to O, N, F called an H-bridge donor and an electronegative O, N, F atom called an H-bridge recipient.





V.2.2 Van Der Waals bond

Furthermore, between molecules, there are weaker bonds such as Van der Waals bonds which ensure the cohesion of liquids and solids. These are low intensity electrostatic type interactions, which allow the cohesion of two dipoles (+ pole attracts towards the - pole).

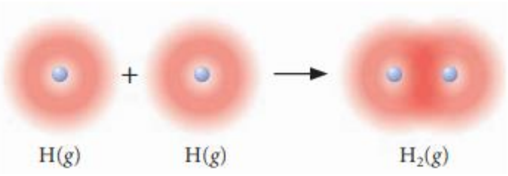


V.2.2. Strong bonds

V.2.2.1. The covalent bond

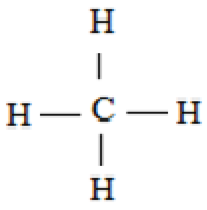
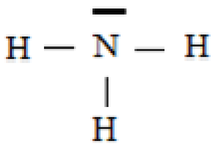
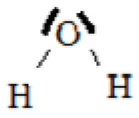
Bond between atoms of neighboring electronegativities. It connects the atoms of non-metals.

Through bonds, atoms generally tend to take the external electronic configuration of a noble gas (2 or 8 valence electrons). In the molecule, the bonds are established so that the symmetry of the molecule is maximum and its minimum energy. There are different types of covalent bond.



a) Simple covalent bond

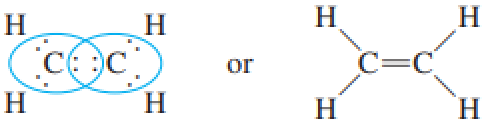
When the covalent bond that links two atoms is made up of only one doublet, we say that it is a single covalent bond. The molecules CH_4, NH_3 and H_2O contain only single bonds:

b) Multiple covalent bond

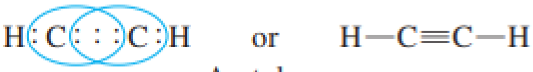
In the molecules we have described so far, each of the bonds has been a single bond that is, a covalent bond in which a single pair of electrons is shared by two atoms. But it is possible for atoms to share two or more electron pairs.

A double bond is a covalent bond in which two pairs of electrons are shared by two atoms. As examples, consider ethylene, C_2H_4. Its Lewis formulas is:



Ethylene

A triple bond is a covalent bond in which three pairs of electrons are shared by two atoms. As examples, consider acetylene C_2H_2. Its Lewis formulas is:

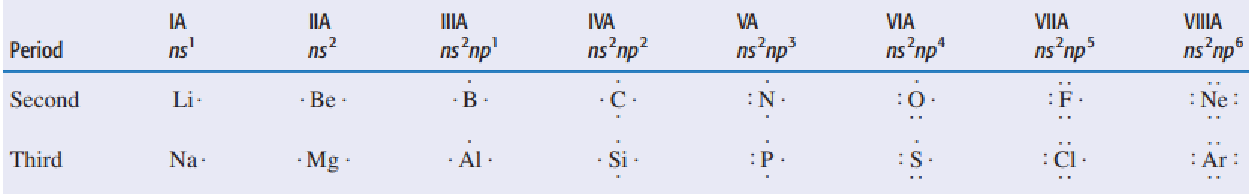


Acetylene

V.2.3 Lewis Electron-Dot Symbols

You can simplify the preceding equation for the electron transfer between Na and Cl by writing Lewis's electron-dot symbols for the atoms and monatomic ions. A Lewis electron-dot symbol is a symbol in which the electrons in the valence shell of an atom or ion are represented by dots placed around the letter symbol of the element.

The table below lists Lewis symbols and corresponding valence-shell electron configurations for the atoms of the second and third periods. Note that dots are placed one to each side of a letter symbol until all four sides are occupied.



Period	IA ns^1	IIA ns^2	IIIA ns^2np^1	IVA ns^2np^2	VA ns^2np^3	VIA ns^2np^4	VIIA ns^2np^5	VIIIA ns^2np^6
Second	$\mathrm{Li}\cdot$	$\cdot\mathrm{Be}\cdot$	$\cdot\overset{\cdot}{\mathrm{B}}\cdot$	$\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{C}}}\cdot$	$:\overset{\cdot}{\underset{\cdot}{\mathrm{N}}}\cdot$	$:\overset{\cdot}{\underset{\cdot\cdot}{\mathrm{O}}}\cdot$	$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{F}}}\cdot$	$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Ne}}}:$
Third	$\mathrm{Na}\cdot$	$\cdot\mathrm{Mg}\cdot$	$\cdot\overset{\cdot}{\mathrm{Al}}\cdot$	$\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{Si}}}\cdot$	$:\overset{\cdot}{\underset{\cdot}{\mathrm{P}}}\cdot$	$:\overset{\cdot}{\underset{\cdot\cdot}{\mathrm{S}}}\cdot$	$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}\cdot$	$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Ar}}}:$

Then the dots are written two to a side until all valence electrons are accounted for. The exact placement of the single dots is immaterial. For example, the single dot in the Lewis symbol for chlorine can be written on any one of the four sides.

The equation representing the transfer of an electron from the sodium atom to the chlorine atom is:

$$\mathrm{Na}\cdot + \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}: \longrightarrow \mathrm{Na}^+ + [:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}:]^-$$

V.2.4 Octet Rule

In each of the molecules we have described so far, the atoms have obtained noble gas configurations through the sharing of electrons. Atoms other than hydrogen have obtained eight electrons in their valence shells; hydrogen atoms have obtained two. The tendency of atoms in molecules to have eight electrons in their valence shells (two for hydrogen atoms) is known as the octet rule. Many of the molecules we will discuss follow the octet rule. Some do not.

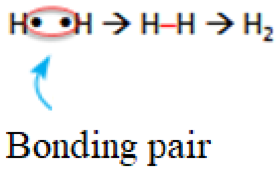
Lewis's structure of diatomic molecules:

- Find the number of valence electrons in each atom.

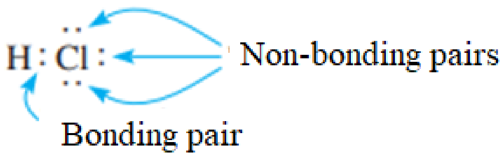
- Place the electrons in such a way as to ensure its octet for each of the atoms except for the H atom which will be surrounded by a doublet.

Examples:

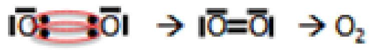
In the dihydrogen molecule, there is a covalent bond between the two H atoms which results from the sharing of a pair of electrons.



The HCl molecule has electron doublets which are not engaged in bonding, these are free doublets, free pairs or non-bonding doublets.

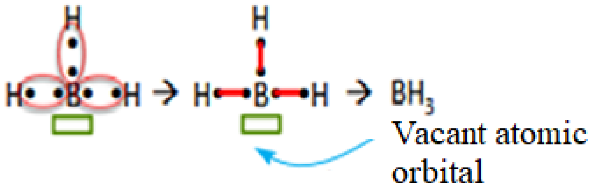


Double bonds form between oxygen atoms:



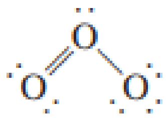
Formation of a triple bond between the two nitrogen atoms:

Presence of vacant atomic orbital.



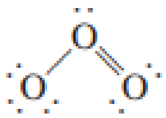
V.2.5 Delocalized Bonding: Resonance

We have assumed up to know that the bonding electrons are localized in the region between two atoms. In some cases, however, this assumption does not fit the experimental data. Suppose, for example, that you try to write an electron-dot formula for ozone, O_3. You find that you can write two formulas:



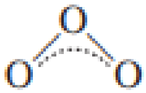
A

and



B

In formula A, the oxygen–oxygen bond on the left is a double bond and the oxygen–oxygen bond on the right is a single bond. In formula B, the situation is just the opposite. Experiment shows, however, that both bonds in O_3 are identical. Therefore, neither formula can be correct. According to theory, one of the bonding pairs in ozone is spread over the region of all three atoms rather than associated with a particular oxygen–oxygen bond. This is called delocalized bonding, a type of bonding in which a bonding pair of electrons is spread over a number of atoms rather than localized between two. We might symbolically describe the delocalized bonding in ozone as follows:



The broken line indicates a bonding pair of electrons that spans three nuclei rather than only two. In effect, the oxygen– oxygen bond is neither a single bond nor a double bond but an intermediate type.

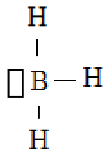
The electron structure of ozone can be described in terms of the two resonance formulas presented at the start of this section. By convention, we usually write all of the resonance formulas and connect them by double-headed arrows.

V.2.6 Exceptions to the Octet Rule

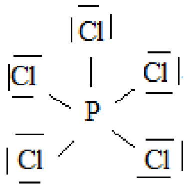
Many molecules composed of atoms of the main-group elements have electronic structures that satisfy the octet rule, but a number of them do not.

A few molecules, such as NO, have an odd number of electrons and so cannot satisfy the octet rule.

Other exceptions to the octet rule fall into two groups: one a group of molecules with an atom having fewer than eight valence electrons around it:



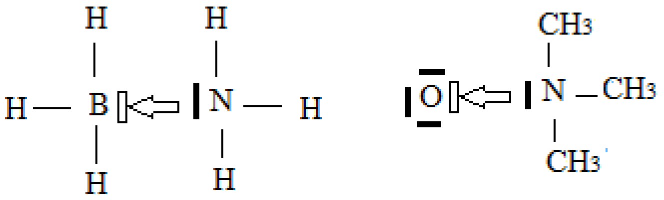
And the other a group of molecules with an atom having more than eight valence electrons around it. The exceptions in which the central atom has more than eight valence electrons around it are fairly numerous. Phosphorus pentachloride (PCl_5) is a simple example. Each molecule consists of a phosphorus atom surrounded by the more electronegative fluorine atoms:



The Lewis model does not provide information on the geometry of the compounds.

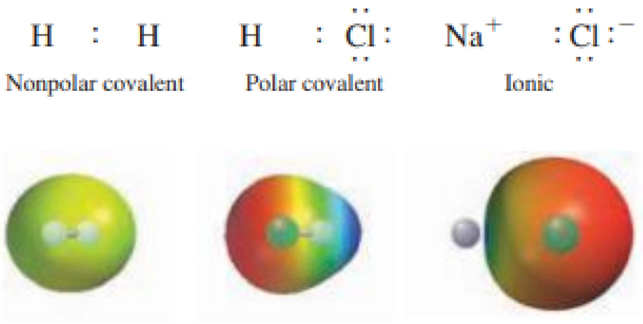
V.2.7 Coordinate covalent bond

When an atom in a molecule does not satisfy the octet, it can bond to another molecule and form a new bond that will cause it to meet this octet rule. In this case, a coordinate covalent bond can be formed provided that one of the atoms plays the role of donor and the other atom plays the role of acceptor of the doublet of the bond.



V.2.8 Polar Covalent Bonds (Electronegativity)

A covalent bond involves the sharing of at least one pair of electrons between two atoms. When the atoms are alike, as in the case of the HOH bond of H_2, the bonding electrons are shared equally. That is, the electrons spend the same amount of time in the vicinity of each atom. But when the two atoms are of different elements, the bonding electrons need not be shared equally. A polar covalent bond (or simply polar bond) is a covalent bond in which the bonding electrons spend more time near one atom than the other. For example, in the case of the HCl molecule, the bonding electrons spend more time near the chlorine atom than the hydrogen atom. You can consider the polar covalent bond as intermediate between a nonpolar covalent bond, as in H_2, and an ionic bond, as in NaCl. From this point of view, an ionic bond is simply an extreme example of a polar covalent bond. To illustrate, we can represent the bonding in H_2, HCl, and NaCl with electron-dot formulas as follows:



The more electronegative atom has a negative partial charge (- δe) while the second, less electronegative atom has a positive partial charge (+ δe).

The difference in electronegativity (χ) makes it possible to determine the nature of the bonds between the atoms.

- $\Delta\chi > 1.7$ The bond is ionic.
- $\Delta\chi < 1.7$ The bond is covalent and polar in nature.
- $\Delta\chi = 1.7$ The bond is 50% polar character and 50% ionic character.
- $\Delta\chi = 0$ The bond is purely covalent.

V.2.9 Ionic bond

It is a bond between atoms with very different electronegativities ($\Delta\chi > 1.70$) It associates metals with non-metals. The metal donates one or more electrons to form a positively charged ion (cation). The non-metal captures these electrons to form a negatively charged ion (anion). The two ions formed have the noble gas configuration. The bond results from the attraction between the cation and the anion.

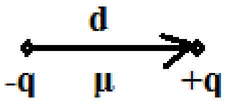
$$Na \longrightarrow Na^{+} + 1e^{-}$$

$$Cl + 1e^{-} \longrightarrow Cl^{-}$$

V.2.10 Dipolar moment

A molecule is polar when it has a dipole moment. In the case of poly-atomic molecules, we add the different dipole moments of all the bonds (sum can be zero in certain cases).

Let two charges –q and +q separated by a distance d generate a dipole moment such that: $\mu = q \times d$



The dipole moment is oriented from the atom carrying the charge (-) to the one carrying the charge (+).

The moment expressed in (Coulomb x meter) and we can use the Debye:

$$1D = 3.34.10^{-30}\ C.m$$

V.3 Molecular Geometry and Chemical Bonding Theory

The Valence Shell Electron Paire Repulsion (VSEPR) model predicts the shapes of molecules and ions by assuming that the valence-shell electron pairs are arranged about each atom so that electron pairs are kept as far away from one another as possible, thus minimizing electron-pair repulsions.

The molecule is symbolized by:

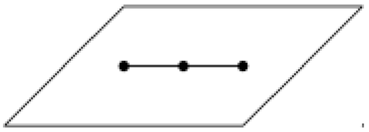
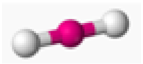
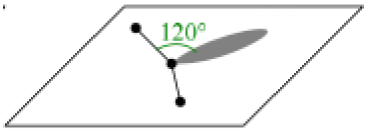
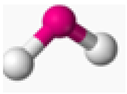
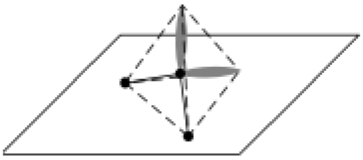
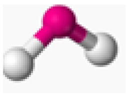
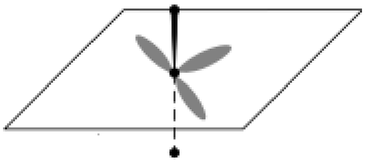
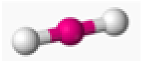
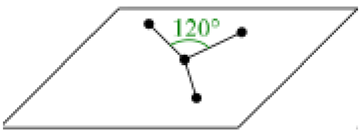
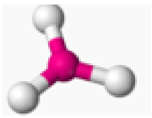
$$A\,X_n\,E_p$$

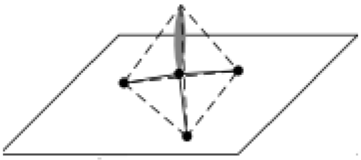
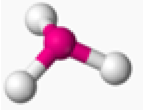
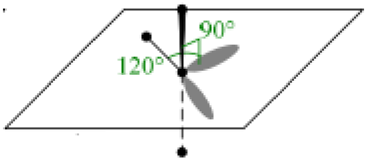
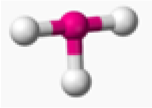
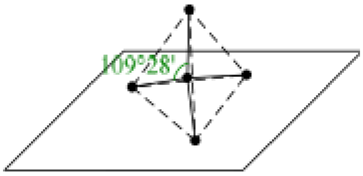
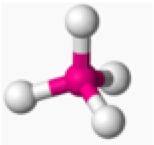
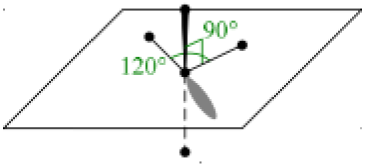
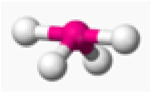
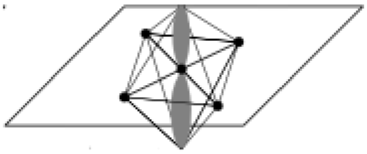
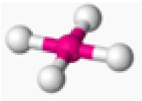
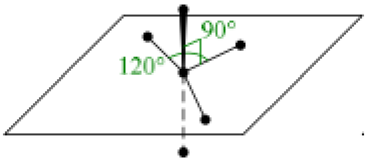
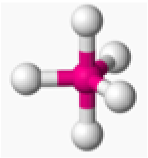
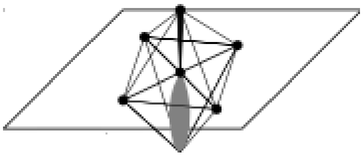
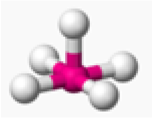
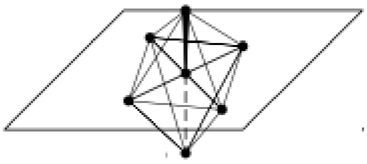
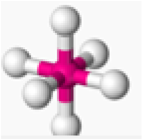
A denotes the central atom which is bonded to n atoms X and which has p pairs of free electrons E or p non-bonding doublets.

Knowing that the electronic doublets of the valence layer repel each other, the geometry of the assembly will be that for which the repulsions are minimal, therefore the electronic doublets are as far apart as possible. We can then predict, from a Lewis diagram, the geometry of the molecule.

(n + p)	Geometry
2	Linear
3	Trigonal planar
4	Tetrahedral
5	Trigonal bipyramidal
6	Octahedral
7	Bipyramid with pentagonal base

Table II.1: Geometry of molecules

Molecule type	Arrangement of Pair	Geometry of the molecule	Shape of the molecule	Examples
AX_2E_0	Linear			$BeCl_2$, $HgCl_2$, CO_2
AX_2E_1	Bent (or angular)			NO_2^-, SO_2, O_3, CCl_2
AX_2E_2	Bent (or angular)			H_2O, OF_2
AX_2E_3	Linear			XeF_2, I_3^-, $XeCl_2$
AX_3E_0	Trigonal planar			BF_3 ,CO_3^- , NO_3^- , SO_3

AX_3E_1	Pyramide A base Triangulaire			NH_3 , PCl_3
AX_3E_2	T shape			ClF_3 , BrF_3
AX_4E_0	Tetrahedral			CH_4, PO_4^{-3} $SO_4^{-2,}$ ClO_4^{-}
AX_4E_1	Non-regular tetrahedron			SF_4
AX_4E_2	square			XeF_4
AX_5E_0	Bipyramid trigonal			PCl_5
AX_5E_1	Pyramid at square base			ClF_5 , BrF_5, $XeOF_4$
AX_6E_0	Octahedral Bipyramid with square base			SF_6, WCl_6

Chapter VI: Organic chemistry

VI.1. Introduction

Organic chemistry is the chemistry of carbon and hydrogen compounds. These compounds also contain heteroatoms such as oxygen, nitrogen, halogen, phosphorus, sulfur, metals and non-metals.

Organic chemistry is the source of almost half of global industrial activity (manufacture of plastic, glass, glues, textiles, perfumes, dyes). Furthermore, organic chemistry constitutes an essential tool for the biologist, it allows him to better understand the phenomena in biochemistry, as well as in pharmacology and agri-food.

VI.2. Characterization of an organic compound

An organic compound is characterized by a chemical formula which can be crude, expanded, semi-developed, compact and topological.

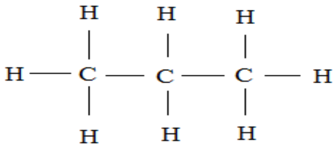
VI.2.1. Crude formula of an organic compound.

The crude formula of an organic compound is made up of x carbon atoms, y hydrogen atoms and z oxygen atoms and is written as follows: $C_x H_y O_{z:}$ C_4H_8, $C_8H_6O_2$

This formula is insufficient to define a compound. It does not specify according to which sequence the atoms are linked.

VI.2.2. Expanded Structural Formulas

In expanded structural formulas, all atoms and bonds are shown figure VI.1.



Figures VI.1 : Formule développée du propane.

VI.2.3. Condensed Structural Formulas

In condensed structural formulas, only specific bonds are shown; this is useful in reducing the number of C—H bonds that must be drawn:

$$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_3$$

Figures VI.2: Condensed Structural Formulas of butane.

VI.2.4. Compact formula

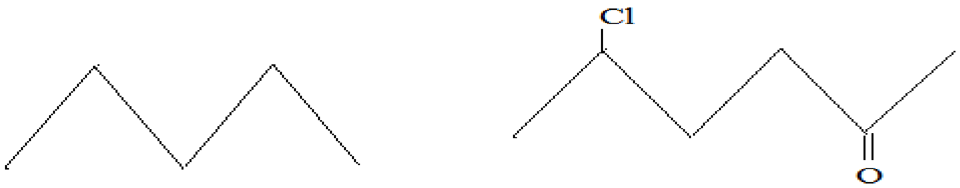
The compact formula of an organic compound simplifies the writing of structural formulas, the bonds are not represented and identical groups are put in parentheses with a subscript which indicates their multiplicity.

$$CH_3(CH_2)_3CH_3$$

Figures VI.3: Compact Formula of pentane.

VI.2.5. Line Drawings

In line drawings (line-angle formulas, skeletal structures, stick figures), bonds are represented by lines; everywhere two lines meet or a line begins or ends is a C atom. H's on C's are not shown (except for emphasis); H's on other atoms must be shown.



(a) (b)

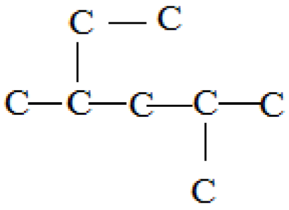
Figure VI.4: Line Drawings of (a) pentane et (b) 5-chlorohexan-2-one.

VI.2.6. Skeleton

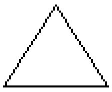
VI.2.6.1. Linear chains

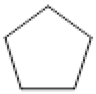
C—C—C—C—C

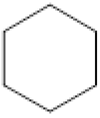
VI.2.6.2. Branched chains

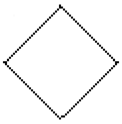


VI.2.6.3. Cyclic chains









VI.3. Nomenclature

In order to name chemical entities, very precise rules have been given by an international organization called IUPAC (International Union of Pure and Applied Chemistry).

VI.3.1. Alkanes

Are acyclic saturated hydrocarbons their nomenclature is: prefix corresponding to the carbon number of the chain plus the ending ane of crude formula C_nH_{2n+2}.

Number of C's	Prefix	Number of C's	Prefix
1	Meth-	7	Hept-
2	Eth-	8	Oct-
3	Prop-	9	Non-
4	But-	10	Dec-
5	Pent-	11	Undec-
6	Hex-	12	Dodec-

Example: CH_3-CH_2-CH_2-CH_3: But + termination ane = Butane

VI.3.1.1. Acyclic branched alkanes

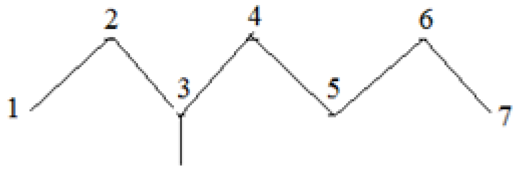
- The main chain is the one with the greatest number of carbons (it is not always written in a straight line).
- The indices indicating the position of the substituents must be as small as possible.
- The name of the substituents does not take **ane** but the ending **yle**.
- Substituent names are placed before the main group.
- If there are multiple substituents, they will be placed in alphabetical order.
- If there are identical substituents: **2 = Di**, **3 = Tri**, **4 = Tetra**…, do not enter the alphabetical order. The terms **iso, sec, tertio, cyclo**, etc., come in alphabetical order.
- If there are several chains of equal length capable of constituting the main chain, we choose:

a) The one with the greatest number of substituents.

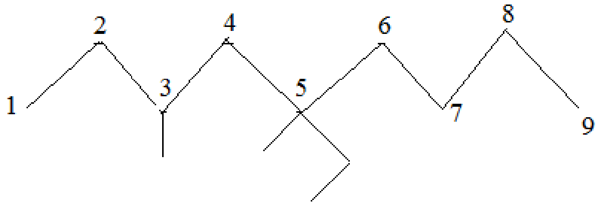
b) When two or more chains satisfy condition (a), we choose the one for which the substituents have the smallest indices.

c) When two or more chains satisfy condition (b), we choose the one for which the lowest indices are given to the substituents named first in alphabetical order.

Example :



3-Methylheptane



5-ethyl-3,5-dimethylnonane

VI.3.1.2. Nomenclature of complex or branched radicals

The principle of the nomenclature of complex or branched radicals is identical to that used in the case of alkanes with linear substituents, the only difference lies in the numbering.

1) The carbon atom with free valence is numbered 1'.

2) The name of the complex radical is considered from the first letter, that is to say the terms Di, Tri, Tetra…, enter in alphabetical order.

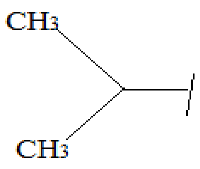
3) In the case where the name of complex radicals are composed of identical words, citation priority is given to the radical which has the lowest position index in its skeleton on the occasion of the first difference.

4) To indicate the multiplicity of complex (branched) radicals, we use the terms **Bis, Tris, Tetrakis, Pentakis,** etc., these prefixes do not fit into alphabetical order.

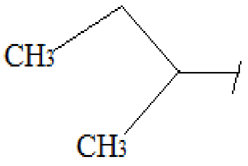
5) The name of the complex radical is put in parentheses preceded by a position index.

VI.3.1.3. Common Substituents

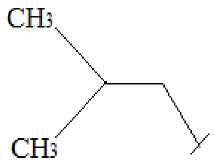
Isopropyl



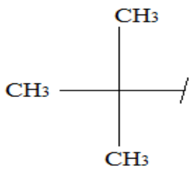
Secbutyl



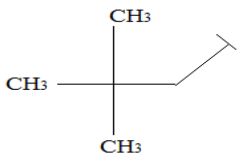
Isobutyl



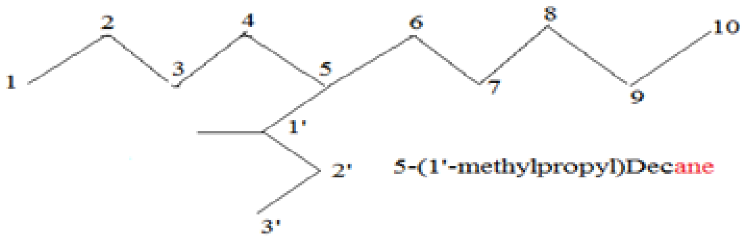
Tertiobutyl



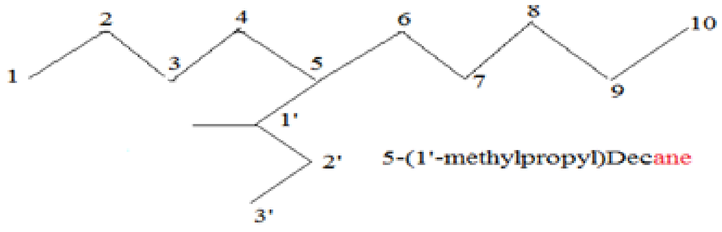
Neopentyl



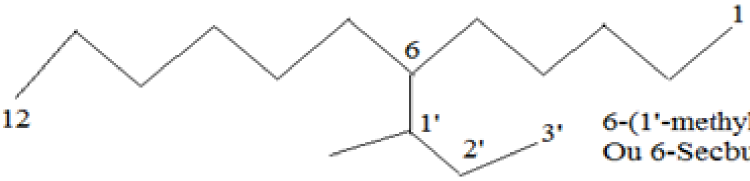
Examples:



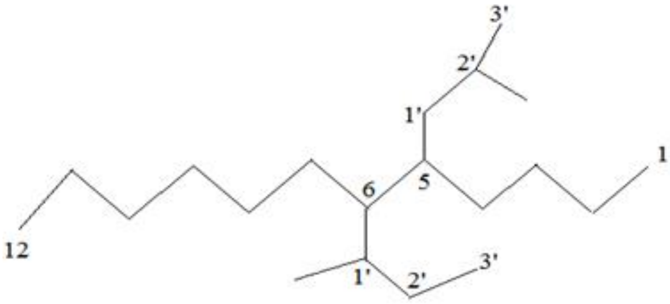
5-(1'-methylpropyl)Decane



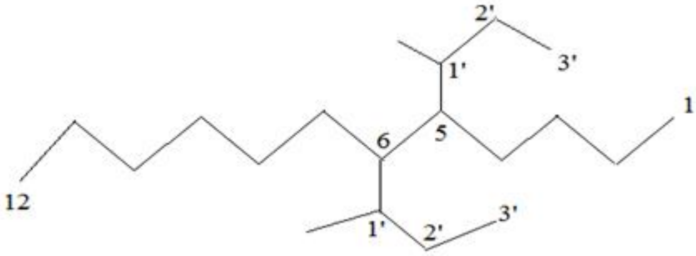
5-(1'-methylpropyl)Decane



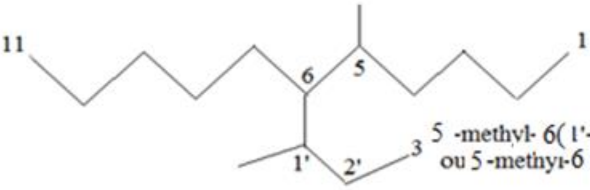
6-(1'-methylpropyl)dodecane
Ou 6-Secbutyldodecane



6-(1'-methylpropyl)-5-(2'-méthylpropyl)dodecane
Ou 5-isobutyle-6-secbutyl dodecane



5,6-Bis-(1'-methylpropyl)dodecane
Ou 5,6-Bis-secbutyl dodecane

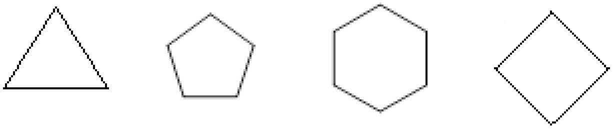


5 -methyl- 6(1'-methylpropyl) undecane
ou 5 -methyl-6 sebutyl undecane

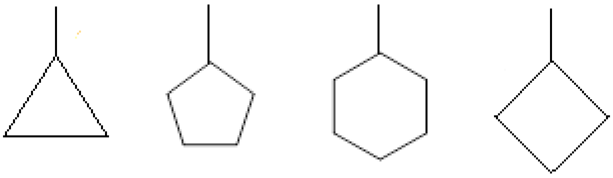
Take notice: According to alphabetical order, the methyl terms are identical, however the methyl substituent takes priority over methyl propyl.

VI.3.2. Cyclic alkanes

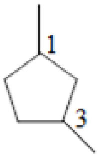
Cyclic alkanes are simply designated by the prefix cyclo followed by the name of the linear alkane corresponding to the numbers of carbons in the cycle: cyclopropane, cyclopentane, cyclohexane, etc.



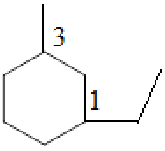
When the cycloalkane is a radical it becomes: cyclopropyl, cyclopentyl, cyclohexyl…



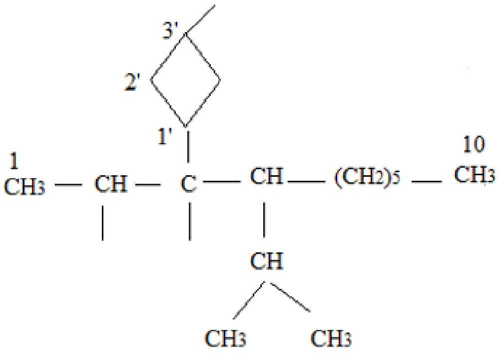
Examples:



1,3-Dimethyl cyclopentane



1-ethyl-3-methylcyclohexane



4-Isopropyl-2,3-Dimethyl-3-(3'-methylcyclobutyl)Décane

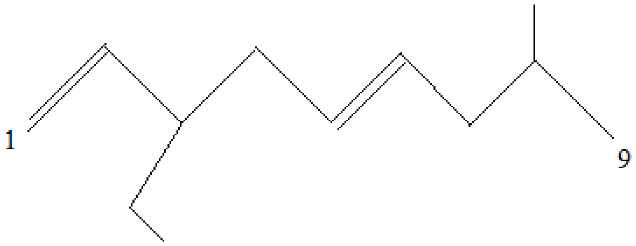
VI.3.3. Acyclic unsaturated compounds

VI.3.3.1. Alkenes

The name of an unsaturated compound with a double bond of the chemical formula C_nH_{2n} is formed by the prefix of the saturated compound with the change of the ending **ane** to **ene**. If there are several double bonds the suffix éne is preceded by the position index of the double bond then by **Di, Tri, Tétra**…

Dial direction is on the side closest to the double bond.

Example :



3-ethyl 8-methyl Nona 1,5-diene

VI.3.3.2. Radicals from alkenes

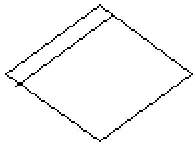
The name of the radical is obtained by replacing the ending **ene** with **enyl**.

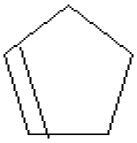
$CH_2=CH-$ ethenyl

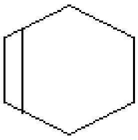
$CH_3-CH=CH-$ propenyl.

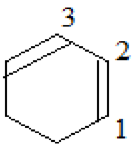
VI.3.3.3. Unsaturated monocyclic compounds

Designated by the prefix **cyclo** followed by the name of the **alkene**.


Cyclobutene


Cyclopentene

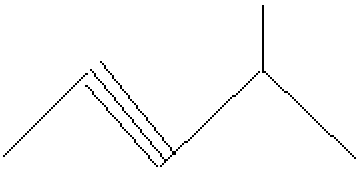

Cyclohexene



Cyclohex-1,3-diene

VI.3.3.4. Alkynes

The name of an unsaturated compound with a triple bond of the chemical formula C_nH_{2n-2} is formed by the prefix of the saturated compound with the change of the ending **ane** to **yne**. The rules of nomenclature are identical to those used for alkenes, except that the position of the triple bond dictates the direction of numbering. If there are several triple bonds the suffix yne is preceded by the position index of the triple bonds then by **Di, Tri, Tetra…**


4-methyl pent-2-yne

VI.3.3.5. Radicals from alkynes

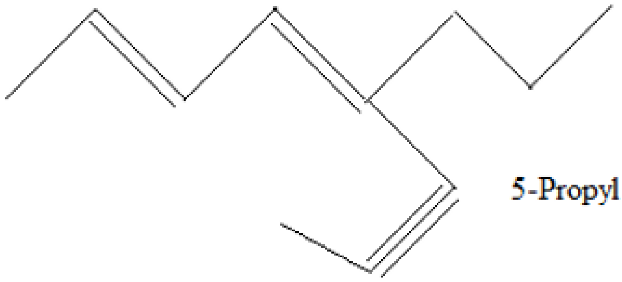
The name of the radical is obtained by replacing the ending **yne** with **ynyl**.

$CH\equiv C-$ Ethynyl

$CH_3-C\equiv C-$ propynyl

VI.3.3.6. Compounds with double and triple bonds

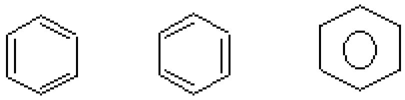
The alkene function has priority over the alkyne function. The main chain contains the maximum number of double and triple bonds.



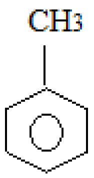
5-Propyl Octa-2,4-diene-6-yne

VI.3.3.3. Monocyclic aromatic compounds

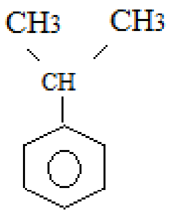
A cyclic compound is aromatic when it has alternating double bonds. The simplest being benzene C_6H_6.



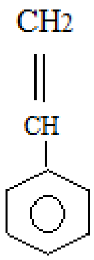
Trade names of benzene



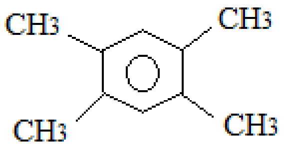
Methyl benzene
Toluene



Isopropyle Benzene
Cumene

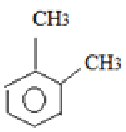


Styrene

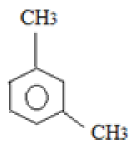


Durene

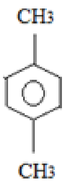
If the benzene ring has two substituents, we can use the terminology ortho (o), meta (m) and para (p) instead of the indices.



O-methyl methyl benzene

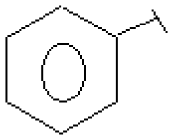


m-methyl methylbenzene

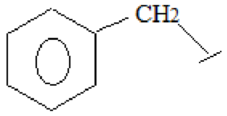


p-methyl methylbenzene

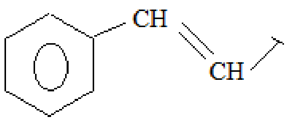
The most commonly used radicals derived from benzene hydrocarbons are:



Phenyl



Benzyl

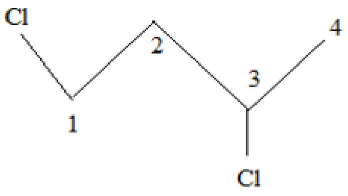


Styryl

VI.3.4. Nomenclature of organic functions

VI.3.4.1. Halogenated derivatives

The halogens are fluorine F, chlorine Cl, bromine Br and iodine I, they are designated by the symbol fluoro, chloro, bromo and iodo themselves preceded by position indices.



1,3-Dichloro Butane

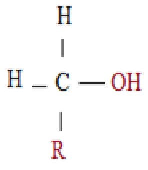
$F_3C — CF_3$

Perfluoro ethane

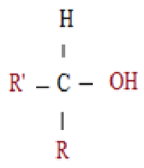
The prefix (per) indicates that all hydrogens in the molecule are replaced by a halogen.

VI.3.4.2. Alcohols

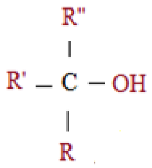
Alcohols have the general formula **R-OH**, OH is the hydroxyl group. They are named by following the name of the corresponding hydrocarbon with the ending **ol** preceded by the position index of the function which must be as low as possible. If the molecule contains several alcohol functions, the ending will be (diols, triols, etc.). The degree of substitution of the carbon carrying the OH group makes it possible to distinguish three classes of alcohols.



$$H-\underset{R}{\overset{H}{C}}-OH$$



$$R'-\underset{R}{\overset{H}{C}}-OH$$

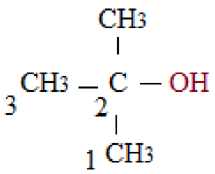


$$R'-\underset{R}{\overset{R''}{C}}-OH$$

Primary alcohol Secondary alcohol Tertiary alcohol

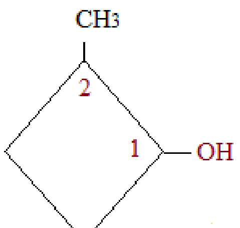
If the alcohol function is not a priority it is designated by the prefix hydroxy.

Examples:

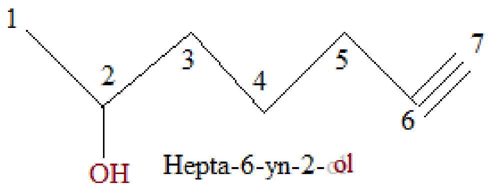


$$\underset{3}{CH_3}-\underset{1\,CH_3}{\overset{CH_3}{\underset{2}{C}}}-OH$$

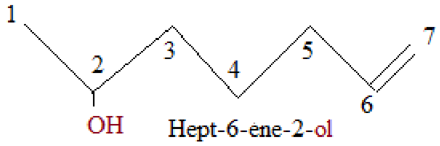
2-methylpropan -2 ol



2-methyl cyclobutanol



Hepta-6-yn-2-ol



Hept-6-ene-2-ol

VI.3.4.3. Ether oxides

The ethers have the general formula R-O-R (symmetric) or R-O-R’ (mixed). To name them, we determine a base name which is that of the hydrocarbon corresponding to the priority R group, that is to say the group which contains the greatest number of carbons or an unsaturation or a function. The RO group-alkoxy group which contains the other R is considered a substituent.

$H_3C\text{-}O\text{-}CH_2\text{-}CH_3$	$H_3C\text{-}CH_2\text{-}O\text{-}CH{=}CH\text{-}CH_3$
Methoxy ethane	Ethoxy propene

We can also name oxide ethers by preceding the names of the groups R and R' with the term oxide...Alkyl groups are named in alphabetical order and this name is mainly used for symmetrical ethers.

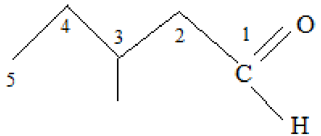
$H_3C\text{-}O\text{-}CH{=}CH_2$	$H_3C\text{-}H_2C\text{-}O\text{-}CH_2\text{-}CH_3$
Methyl and vinyl oxide	Diethyl oxide

VI.3.4.4. Aldehydes

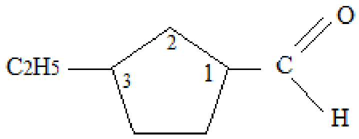
Aldehydes have the general formula (R-CHO).

a) The name of an acyclic aldehyde is formed by adding the ending **al** (or **dial, trial**, etc.) to the name of the corresponding hydrocarbon. The carbon in the **CHO** group is numbered 1 and the function position index is usually omitted.

b) Cyclic aldehydes in which the function is directly linked to the ring, are named by adding the ending carbaldehyde to the name of the ring.

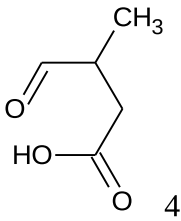


3-methyl pentanal



3-ethyl cyclopentane carbaldehyde

c) If the aldehyde function is not a priority it is designated by the prefix formyl.



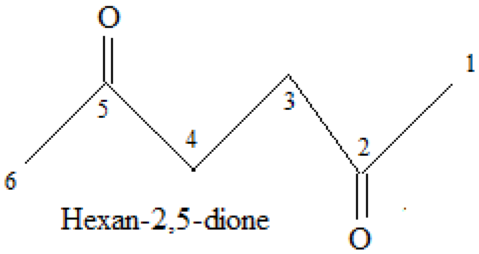
4-formyl-3-methyl butanoic acid

VI.3.4.5. Ketones

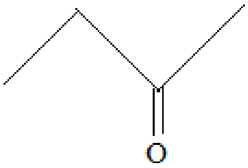
Ketones have the general formula **R-CO-R** (symmetric) or **R-CO-R’** (mixed).

a) The name of a ketone is formed by adding the ending **one**, preceded by the position index, to that of the corresponding hydrocarbon. The main chain is the longest chain that contains the **CO** group.

b) If the groups **R** and **R'** are simple, we can also name the ketones by following their names in alphabetical order with the word ketone.

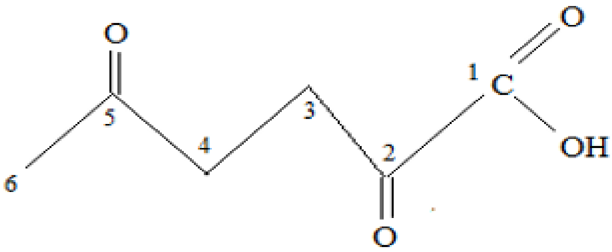


Hexan-2,5-dione



Ethyl methyl Ketone
Butan-2-**one**

c) When the ketone function does not have priority, the **CO** group is then called **oxo**.

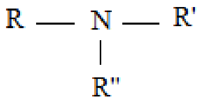


2,5-Di**oxo** Hexanoic acid.

VI.3.4.6. Amines

An amine is an organic compound derived from ammonia **NH_3** in which at least one hydrogen atom is replaced by a carbon group. There are three classes of amines :

$R — NH_2$ $R — NH — R'$



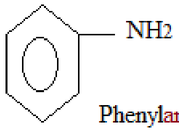
$$\begin{array}{c} R — N — R' \\ | \\ R'' \end{array}$$

Primary amine Secondary amine Tertiary amine

a) Primary amines are named by adding the amine ending to the R group, if R is branched, its main chain must contain the carbon linked to $–NH_2$ and which has the index 1.

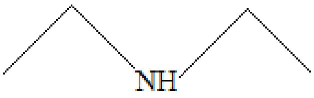
$CH_3 — CH_2 — NH_2$

Ethylamine

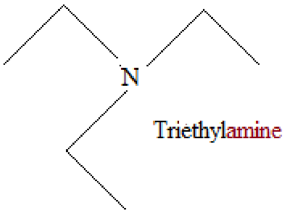


Phenylamine

b) When the secondary and tertiary amines are symmetrical (the **R** groups identical), they are named like the primary amines preceded by the prefixes: **di, tri, tetra,** etc.

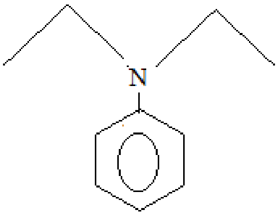


Diethylamine

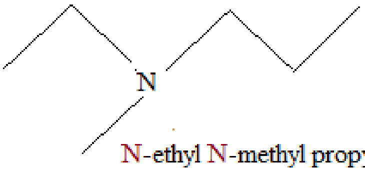


Triethylamine

c) In the case where the amines are mixed (non-identical **R** groups), they are considered as derivatives of the primary amine which has the longest or most complex **R** group. The names of the other groups are placed in front of that of the primary amine by preceding them with the letter **N**.



N.N- diethyl phenylamine

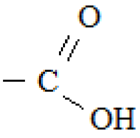


N-ethyl N-methyl propylamine

When the amine function does not have priority, the NH_2 group is then called **amino**.

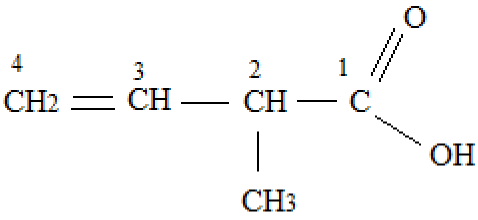
VI.3.4.7. Carboxylic acids

These are organic compounds comprising the carboxyl group linked to an aliphatic group **R** or an aromatic group **Ar**.

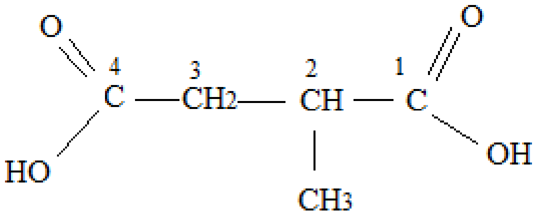


$$-C(=O)OH$$

They are named by following the word acid with the name of the alkane with the same carbon skeleton where the final "e" is replaced by the suffix: **oic**.

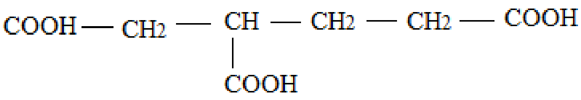


2-methyl but-3-en-1-oïc acid



2-methylbut-1,4-dioïc acid

When, in a polyfunctional compound, several carboxylic groups are distributed over several chains, the prefix **carboxy** is used to designate the function belonging to a branch.

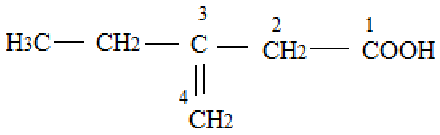


$$COOH-CH_2-CH(COOH)-CH_2-CH_2-COOH$$

3-Carboxyl hexan-1,6-dioic acid

These are compounds that contain the carboxyl group **COOH**.

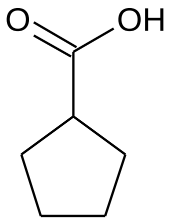
a) In the acyclic series, the acids are named by following the name of the corresponding hydrocarbon with the ending oic (dioic, trioic, etc.), and by ending it with the word acid. The carbon of the **COOH** group always has the index 1, because the acid function has the highest priority Table VI.1.



$$H_3C-CH_2-\overset{3}{C}(=\underset{4}{C}H_2)-\overset{2}{C}H_2-\overset{1}{C}OOH$$

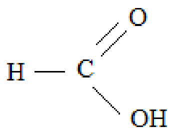
3-ethyl But-3-enoic acid.

b) En série cyclique, les acides dont la fonction est directement liée à un cycle sont nommés en faisant suivre le mot acide du nom de l'hydrocarbure cyclique correspondant auquel est ajouté le suffixe carboxylique.

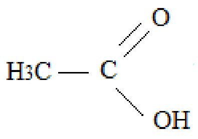


cyclopentanecarboxylic acid

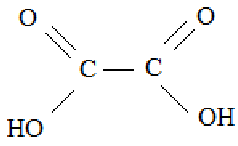
c) Many acids have common names such as:



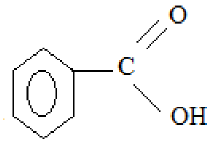
Formic acid



Acetic acid



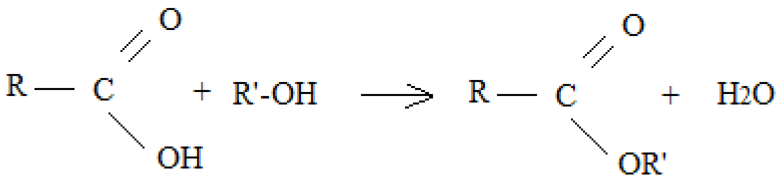
Oxalic acid



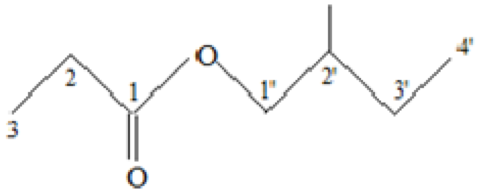
Benzoic acid

VI.3.4.8. Esters

They result from the reaction of acids with alcohols:



Esters have two carbon chains separated by an oxygen atom. The two chains must be named separately; in the name of an ester two terms appear: one an alkanoate, the other an alkyl group.

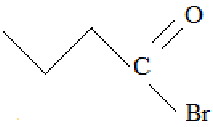


2’-methylbutyl propanoate

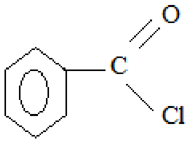
VI.3.4.9. Acid halides

The substitution in a carboxylic acid of the OH group by a halogen X generates an acid halide of general formula R-CO-X.

Acid halides are named by preceding the name of the corresponding hydrocarbon with the term: Halide, and the ending oe becomes oyl.



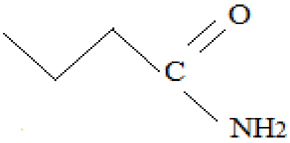
Butanoyl bromide



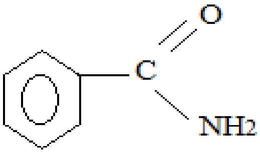
Benzoyl chloride

VI.3.4.10. Amides

The substitution in a carboxylic acid of the **OH** group by an **NH_2** group generates a primary amide **R-CO-NH_2**. Hydrogens can be replaced by alkyl groups.

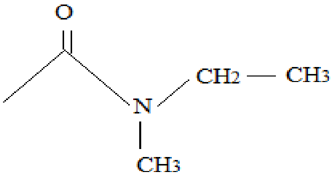


Butanamide

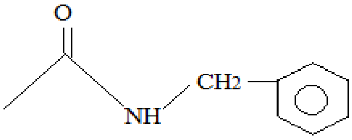


Benzamide

When the amide is substituted R-CO-NR'-R", the names of the alkyl groups are stated before that of the amide by preceding them with the letter N.



N-ethyl-N-methyl ethanamide



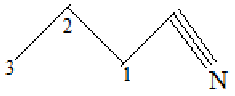
N-benzyl ethanamide

VI.3.4.11. Nitriles

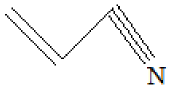
Nitriles have the general formula R-C≡N, they are named by adding the nitrile ending to the name of the corresponding hydrocarbon.

$H_3C\text{-}CH\text{=}CH\text{-}CH_2\text{-}C\equiv N$ pent-3-ene nitrile

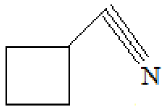
Nitriles can also be named by preceding the name of the radical **R** with the term **cyanide**, but in this case the **C≡N** group is not part of the main chain.



Propyl cyanide



Vinyl cyanide



Cyclo butyl cyanide

When the nitrile function does not have priority, the C≡N group is then called **cyano.** Mixed-function compounds contain different functions in the same molecule. We choose the longest chain which contains the main function. This is designated by a suffix and the direction of numbering of the main chain is chosen in such a way as to give it the smallest possible index, the other functions present are designated by prefixes which are indicated in Table VI. 1.

Table VI.2: Ranking of functions by priority.

Function	Function	**Function Formula**	**Function Formula Prefix**
Carboxylic acid	R-COOH	**Carboxy…**	**Oïc… Acid**
Ester	R-COOR’	**Alkoxycarbonyl…**	**…yle...oate**
Halides acid	R-COOX	**Halocarbonyl**	**oyle… Halogenure**
Amides	R-CO-NH2	**…amido**	**…amide**
Nitriles	R-C≡N	**Cyano…**	**…nitrile**
Aldehydes	R-CHO	**Formyl…**	**…al**
Ketones	R-CO-R	**Oxo…**	**…one**
Alcohols	R-OH	**Hydroxy…**	**…ol**
Amines	R-NH2	**…amino**	**…amine**
oxides	R-O-R’	**Alkoxy…**	**Oxyde of …**
Alkenes	-CH=CH-	**-**	**…ene**
Alkynes	-C≡C-	**-**	**…yne**
Alkanes	CH3-CH2-	**-**	**…ane**
Halides	R-X	**Halogen**	

VI.4. Stereochemistry

Stereochemistry is the field of chemistry that studies the architecture and three-dimensional conformation of molecules, as well as their consequences on physicochemical properties. Compounds that have the same chemical formula but different structural formulas are called "isomers." They differ by the nature of the chemical bonds (planar isomerism) or by the spatial arrangement of the atoms (stereoisomerism).

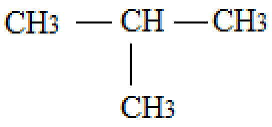
VI.4.1. Plane isomerism

Two planar isomers are isomers that have the same structural formula but different semistructural formulas. These cases of isomerism can be represented by plan diagrams. This isomerism can be of chain, position or function.

VI.4.1.1. Chain isomerism

In this case the sequence of carbons is different, example C4H10

$H_3C\text{-}CH_2\text{-}CH_2\text{-}CH_3$
n-butane



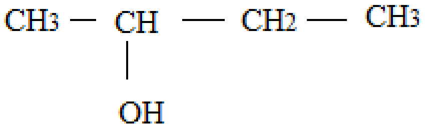
$$CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3$$

2-methyl propane

VI.4.1.2 Position isomerism

In this case we find the same function but occupies different positions:

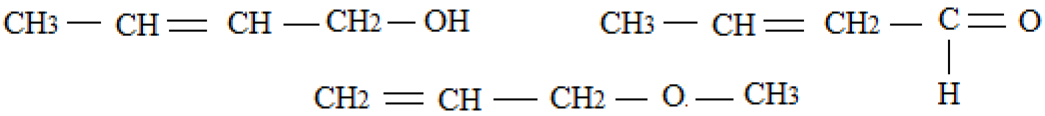
$CH_3-CH_2-CH_2-CH_2-OH$



$$CH_3-\underset{\displaystyle OH}{\underset{|}{CH}}-CH_2-CH_3$$

VI.4.1.3. Functional isomerism

If the isomers differ by the chemical function present in the molecule, we say that they are functional isomers. This is the case for the following three molecules which all have the chemical formula C4H8O. The first is an alcohol, the second an aldehyde and the third an ether oxide.



$CH_3-CH=CH-CH_2-OH$

$$CH_3-CH=CH_2-\underset{\displaystyle H}{\underset{|}{C}}=O$$

$CH_2=CH-CH_2-O-CH_3$

VI.5 Organic reagents

VI.5.1 Electrophilic reagent

It is a positively charged ion (+) capable of forming a new bond by gaining a pair of electrons, it is denoted (E+).

Example: NO_2^+, H_3O^+ Lewis acids have an empty box: BF_3, $CoCl_3$, ZnCl2, $AlCl_3$…

VI.5.2 Nucleophilic reagent

It is a negatively charged ion (-) capable of forming a new bond by sharing two free electrons, it is denoted (Nu-).

Example: OH-, I-, CN-, RO- and the Lewis bases: H2O, ROH, NH3…

VI.5.3 Free Radical Reagent

Resulting from a homolytic breakage of a bond, it is endowed with great reactivity.

Cl-Cl plus UV light gives two Cl• radicals

There are three types of reaction intermediates with opposite charges to the reactants.

Group of atoms having lost an electron, it therefore carries a positive charge (+)

CH3+(nullary carbocation), CRH2+(primary), CR2H+(secondary),CR3+(tertiary).

A tertiary carbocation is more stable than a secondary carbocation which is itself more stable than a primary carbocation

A group of atoms of which one carbon is negatively charged, it is formed by the removal of a proton using a very strong base.

CRH2-(primary), CR2H-(secondary), CR3-(tertiary). A primary carbanion is more stable than a secondary carbocation which is itself more stable than a tertiary carbocation.

VI.5.4 Free radical carbon

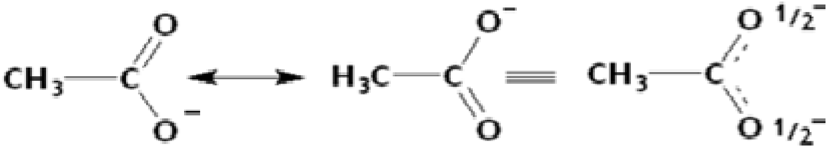
Comes from a homolytic breakage of a bond under the influence of heat or light (UV). CRH2• (primary), CR2H• (secondary), CR3• (tertiary). A tertiary free radical is more stable than a secondary free radical, which is itself more stable than a primary free radical.

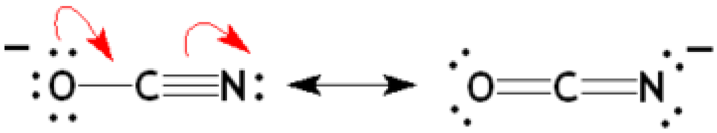
VI.7.3. Resonance and mesomerism

Writing several formulas to represent the same compound constitutes the mesomerism method.

Example The cyanate ion:

Mesomerism in this case explains its reactivity, because this ion can give rise to cyanate and isocyanate compounds by reaction of nitrogen or oxygen, depending on the **conditions.**





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